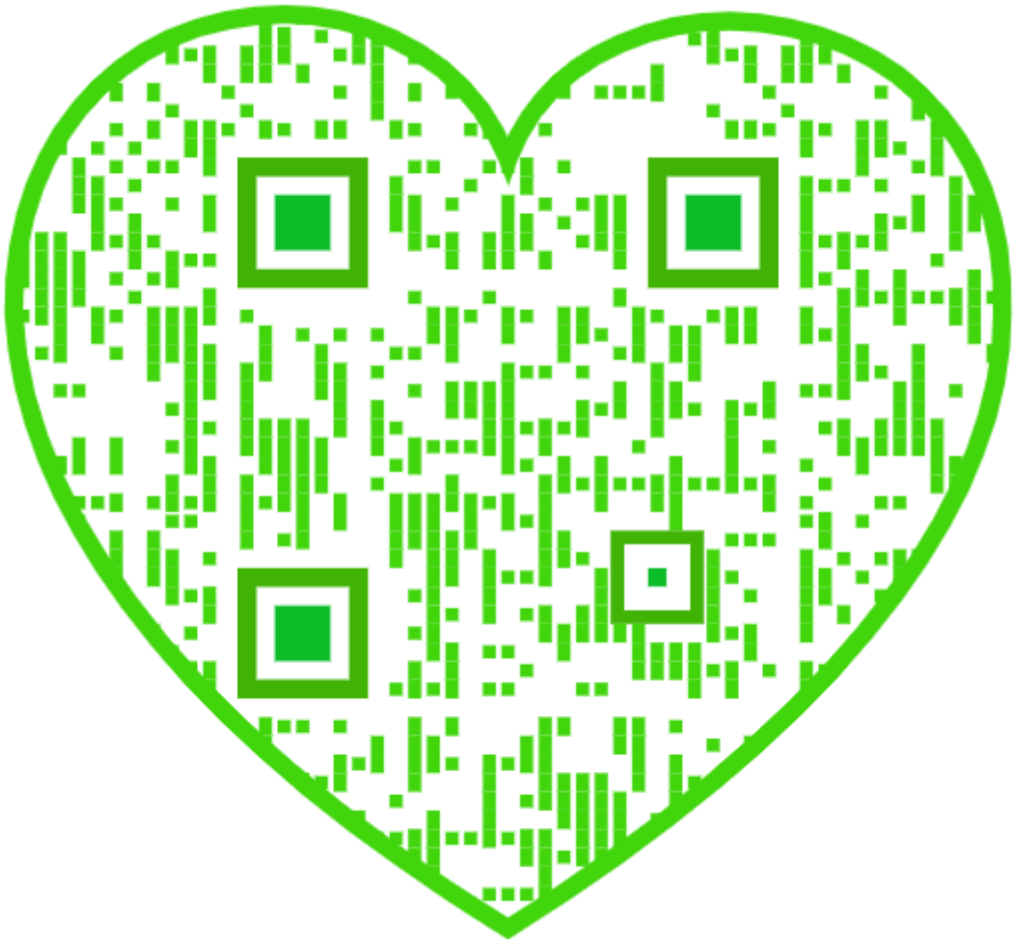

THE

CHEMISTRY OF MEDICINES,

PRACTICAL.

A TEXT AND REFERENCE BOOK

FOR THE USE OF

STUDENTS, PHYSICIANS, AND PHARMACISTS,

EMBODYING THE

PRINCIPLES OF CHEMICAL PHILOSOPHY AND THEIR APPLICATION TO THOSE CHEMICALS THAT ARE USED IN MEDICINE AND IN PHARMACY, INCLUDING ALL THOSE THAT ARE OFFICINAL IN THE PHARMACOPŒIA OF THE UNITED STATES.

With Fifty Original Cuts.

By J. U. LLOYD,

Professor of Chemistry in the Eclectic Medical Institute; Professor of Pharmacy in the Cincinnati College of Pharmacy; Author of "Elixirs," etc., etc., etc.

FOURTH EDITION.

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INTRODUCTION.

It is designed that this work shall include all the organic and inorganic chemicals employed in medicine, and which may be required of pharmacists. These chemicals are practically investigated, as well as the processes by which they are made. The principal impurities or adulterations that may exist in them are named, and proper tests are given for their detection. Poisons are specified, and with each noxious article the treatment and antidotes for injurious doses. Under the head, "synonyms," with each compound will be given the various names by which the article is known to chemists, as well as to commerce. The atomic weights, specific gravities, solubilities, etc., herein given, are the results of the investigations of the most careful scientists. Names of experimenters are seldom given, and this would seem to be an act of injustice towards men who have labored to bring our science to its present condition, but the necessity for condensation, very much to the author's regret, has rendered it necessary for him to omit this act of courtesy, and to simply acknowledge his indebtedness at this place, in general terms. Prof W. L. Dudley kindly reviewed that portion of the manuscript embracing chemical philosophy, for which the

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[illegible] I have endeavored to avoid all unnecessary expressions, and to give an explanation of each reaction in the simplest possible manner consistent with clearness. It has been thought best to give the prominent characters of all elements, that have been indisputably accepted as such, regardless of their use in medicine. In studying compounds, however, I have noticed, as the rule, only such as are used in medicine or by pharmacists, or that are dangerous poisons. The exceptions are certain combinations of non-metallic elements, and metals with oxygen, which are of general interest. In most cases, it will be found that not only are explanations of chemical reactions clearly made in the text, but in addition, equations have usually been introduced as accompaniments. The student

should remember that formulæ and equations only represent compounds and chemical changes; they are simply an aid to the study of facts, and much valuable time may be lost by attempting to memorize unnecessary formulæ and long lines of equations. By introducing only medicinal chemicals, I have been enabled to give, in most instances, quite thorough descriptions of the important ones. My laboratory and commercial experience have been drawn upon to render the work as complete and practical as possible, and my aim has likewise been to make it of interest and value to students, physicians, and druggists. That it may accomplish these purposes is the sincere desire of the author.

J. U. LLOYD.

PREFACE TO SECOND EDITION.

The fact that the first edition of this work was entirely disposed of within a month from the time of its publication, taken in connection with the many complimentary and favorable notices of it from the foremost medical and pharmaceutical journals of our country, as well as from numerous teachers of medicine and of pharmacy, has led me to hope that it may meet with a more extensive circulation, and prove of greater utility, than I could at first reasonably anticipate. I have therefore been induced to present this second and improved edition.

The student should constantly keep in view the fact that the study of medicinal chemicals is simply an application of general, well known chemical laws to such agents and compounds as are employed in medicine. A partial instruction in these laws may be sufficient for the manner of teaching in many schools; but during all such special teaching, and even during the study of any one remedial agent or compound, reference should invariably be made to the general laws L.

THE CHEMISTRY OF MEDICINES.

APPARATUS, WITH DIRECTIONS FOR USE.

In making chemical preparations, and in trying experiments in chemistry, a certain amount of apparatus is necessary. Often the student will purchase a line of apparatus that can not be adapted to the experiments and work which is to be performed. Again, he will labor under disadvantage from not understanding how to handle and use the ware. In view of the above, and to familiarize the student with most of the apparatus mentioned in this work, it has been thought advisable to introduce cuts of the more important pieces with appropriate remarks. It is possible to perform many experiments with the most simple forms of apparatus, and it

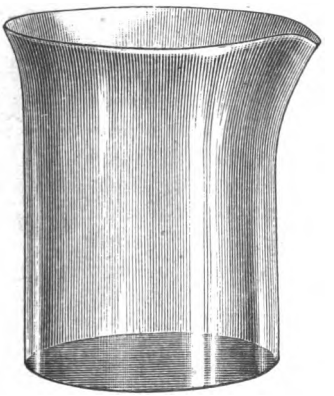

Fig. 1. Beaker Glass.

will be noticed that the directions we give for making most of the compounds in the work to follow require very little beyond a flask or a dish, etc.; therefore the student must not infer that it is necessary for him to purchase the entire outfit at once. Let him consider the experiments which he proposes to try, and from time to time procure the apparatus necessary.

Beaker Glasses.—These are represented by Fig. 1. They are thin, cylindrical glass vessels, and are very useful for mixing liquids, and especially for the purpose of producing precipitates. They are also convenient for filtering into, and as they are made of very thin glass, they will stand the change of temperature caused by hot solutions. They may be purchased singly of any size, from one ounce capacity to half a gallon. A nest, which contains from ten to twelve glasses, one within the other, and holding severally from one ounce to two pints, may be obtained, and this is desirable if many are required. There are two forms of the beaker glass, one *with lip*, as our cut represents, the other is plain. Use the beaker glass with lip.

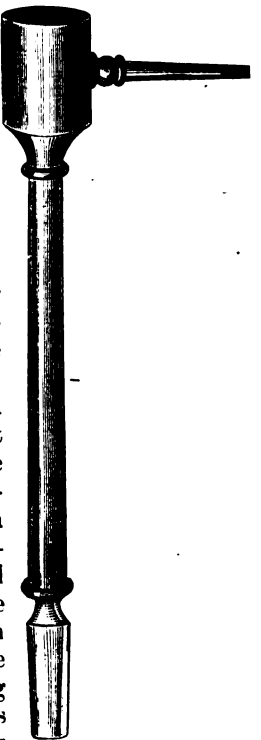

Fig. 2.
Mouth Blowpipe.

Mouth Blowpipe.— This is a very convenient little instrument (fig. 2) in testing for the presence of certain elements, or for the purpose of oxidizing or deoxidizing a substance. The point of the blowpipe is to be inserted into an alcohol flame, or a gas jet, and a gentle current of air then blown through it. This produces a hot double flame having the shape of two long pointed cones, the inner of which is blue, the outer yellow. If a body be placed at the tip of the outer

flame it will be oxidized. If it be placed within this flame and at the tip of the inner flame, deoxidation (reduction) takes place. In consequence of this fact the outer flame is called the *oxidizing* flame, and the inner the *reducing* flame. The blowpipe, as far as experiments in synthesis are concerned, is seldom employed, but it is very useful in analytical operations.

Cork Borers.—These are cylindrical tubes of brass (fig. 3) with a knob upon one end of each, the opposite end being sharpened to an edge. In using this instrument the sharp end is placed against the cork, and by means of a twisting push the tube is thrust through it, thus cutting a cylindrical hole the size of the borer. Cork borers reach the market in sets of from three to twelve, nested and of different sizes. In practice we prefer a rat-tail file for boring holes through corks. With two or three sizes of this file, regular and even holes of any diameter can be quickly drilled.

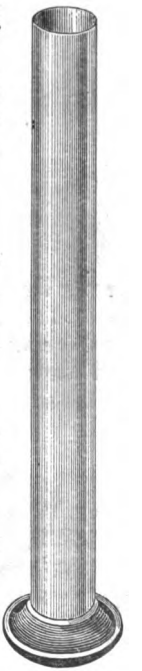

FIG. 3.
Cork Borer.

Crucibles.—Fig. 4 represents the common Hessian (sand) crucible. There are plumbago crucibles and porcelain crucibles in addition, but for ordinary purposes the Hessian crucible will answer, and it is cheapest. Crucibles are used where a high heat is required, as for example, in making monosulphide of calcium (see index). In certain cases *platinum* crucibles are used, but this is not necessary in the practical work of the pharmacist, although for many analytical operations platinum is indispensable.

Evaporating Dishes. — Evaporating dishes are made of porcelain and of glass, but the porcelain dish only is of practical value. Porcelain evaporating dishes are of many qualities, and of different shapes. Our figure 5 represents the most desirable pattern, especially for the large sizes. These have a heavy rim around the top, as shown in the engraving. Small sizes are plain. The

ware known as Berlin porcelain is most reliable, and if properly managed will stand considerable heat. These dishes should be thin and of an even thickness, as may be shown by holding them before a window or other

FIG 4. Hessian Crucible.

strong light. In heating an evaporating dish, if the dish be cold the heat must be gradually raised by means of a sand or a water bath, or by passing the dish backward and forward over a flame until it is warm. Porcelain dishes will stand the action of strong acids, but caustic

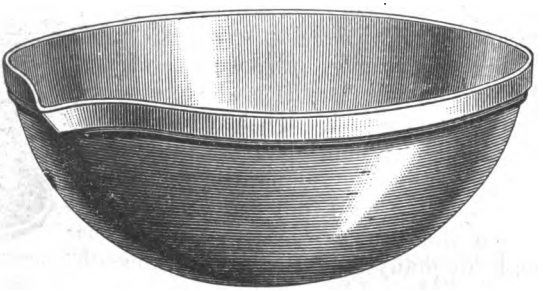

FIG. 5. Evaporating Dish.

alkalies corrode them. They may be obtained of any capacity between one fluid ounce and four gallons. Let it be remembered that in all cases where a "dish" or "basin" is referred to in this work, without further specification, the porcelain evaporating dish is intended. The

term "capacious" is often used by chemists with reference to the vessel that shall be employed. In such cases the vessel should have a capacity of two or three times that of the material it is expected to contain.

Funnel.—This should be of glass, and of the shape represented by our engraving (fig. 6.) Many funnels of the market are unfit for use by reason of the rough, bungling exit tube. Some of these are of thick, heavy glass so large in diameter that they can not enter the mouth of an ordinary bottle, and yet the glass is so thick as to permit the passage of only a very small stream of liquid. In selecting a funnel choose that in which the exit tube is slightly tapering, of light glass, and rather small. The conical part of the funnel should incline at an angle of 60°. It is well to have several sizes of funnels.

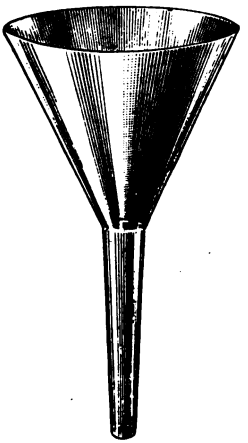

FIG. 6. Funnel.

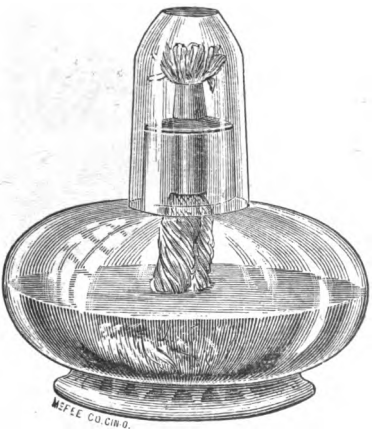

FIG. 7. Alcohol Lamp.

Lamp. — The alcohol lamp (fig. 7) is all that may be required in experiments, but for practical work on a large scale the gas stove is hardly to be dispensed with. Alcohol lamps produce a hot flame at moderate expense, and they are perfectly safe. The glass cover which fits over the wick should always be in place when the lamp is not burning, as it prevents the evaporation of alcohol. Do

not place cork around the tin tube which enters the lamp, as this must fit loosely, in order that air may enter the lamp as the alcohol is used.

Sand Bath.—The sand bath (fig. 8) is simply a thin iron or copper vessel partly filled with sand. The sand bath distributes heat evenly to any vessel that requires a regular temperature. In using the sand bath, place a layer of dry sand half an inch in thickness upon the bottom of the vessel, and upon this place the vessel which is to be heated. Now fill with dry sand the space between

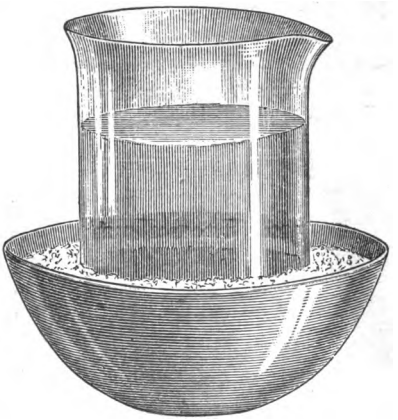

FIG. 8. Sand Bath, with Beaker Glass.

the two vessels. Place the apparatus upon the ring of a filter-stand, and apply heat to the bottom of the outer vessel; or place the arrangement upon a gas stove, or even an ordinary stove. It must be observed that the sand bath should be prepared for use while cold, as hot sand is liable to crack an evaporating dish or a flask.

Water Bath (or expanded steam bath).—This is used upon the same principle as the sand bath and for the same purposes. The water bath has the advantage, however, of an even temperature and a temperature that can not rise above 100° C. (212° F.) The water bath is simply made by placing a pan of water upon the ring of a retort stand and covering the top of the pan with a sheet

of tin through which a hole is cut. The vessel to be heated is placed in position over the opening. It is not necessary that it enter the water, as the heat imparted by

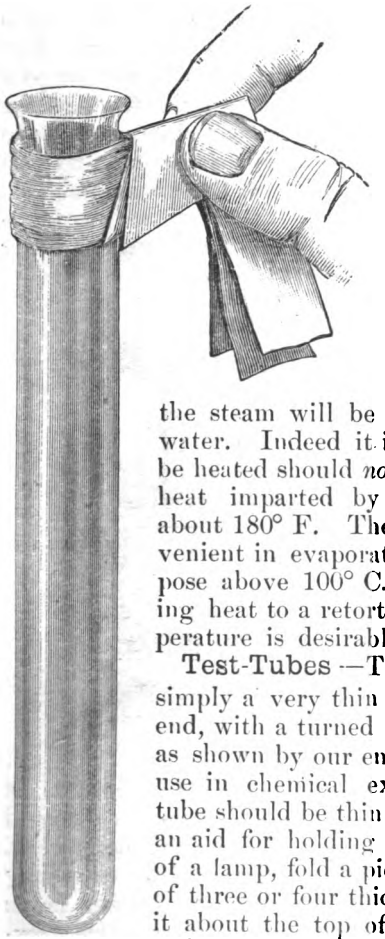

Fig. 9
Test-Tube.

the steam will be as great as that of the water. Indeed it is best that the vessel to be heated should *not* touch the water. The heat imparted by a water bath averages about 180° F. The water bath is very convenient in evaporating liquids that decompose above 100° C. (212° F.), or for applying heat to a retort where a low even temperature is desirable.

Test-Tubes.—The test-tube (fig. 9.) is simply a very thin glass tube closed at one end, with a turned edge upon the opposite, as shown by our engraving. It is of great use in chemical experiments. The test-tube should be thin and of good glass. As an aid for holding the tube in the flame of a lamp, fold a piece of paper into a strip of three or four thicknesses, and then wrap it about the top of the tube, twisting the ends as represented by fig. 9. The operator must always incline the test-tube, when in

use, so that the mouth is directed away from his person. This is to avoid the chance of a spurting liquid which is sometimes thrown from the tube with considerable violence when boiling commences. This is especially liable to occur when the heat is applied at once directly to the bottom of a tube which contains a heavy solution. Experience will teach the student that it is best to first pass the tube backward and forward through the flame of the lamp, and when its contents are warm to incline the tube slightly, with the mouth away from his person. By this means the flame of the lamp heats the entire solution evenly, and boiling begins about the same time throughout the entire liquid. Test-tubes come in nests of four (usually), and are sold at a reasonable rate.

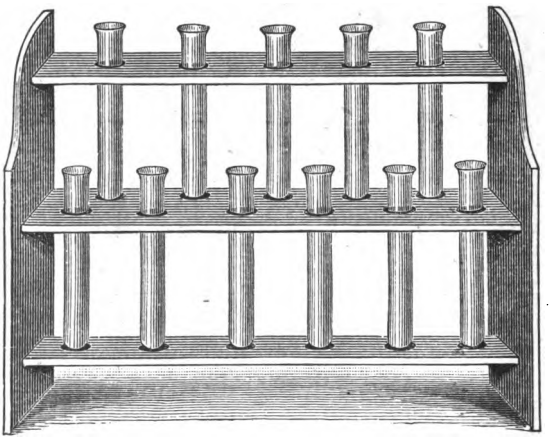

FIG. 10. Test-Tube Rack.

Test-Tube Rack.—This is represented by fig. 10. It is a wooden rack, the holes of each shelf being so arranged that the tubes pass through and rest upon the shelf below. Test-tube racks (or stands) are very convenient, especially where it is desirable that liquids shall stand a few hours for future observation. They are not indispensable, however, and the beginner may employ a row of wide-mouth bottles as test-tube holders with entire satisfaction.

Chemical Flask.— This we represent by fig. 11. The chemical flask is a thin glass vessel, shaped as the cut represents and intended for heating liquids, although it may be used in many other chemical operations This flask is sometimes called the *Florence flask*, a name derived from the fact that some years ago olive oil was imported from Italy in flasks somewhat resembling the foregoing. These were

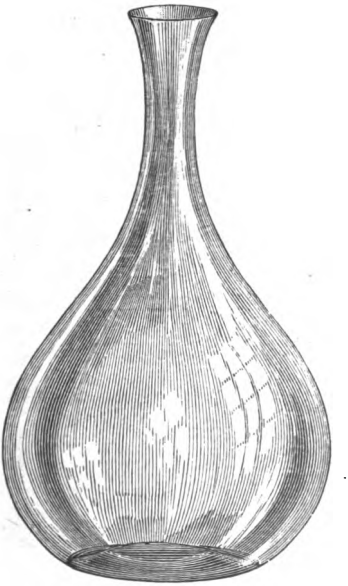

FIG. 11. Chemical Flasks.

employed, when empty, by chemists for the purpose filled by the chemical flask of the present day. Chemical flasks are found in the market with round bottoms and with flattened or concave bottoms. The concave-bottomed flask is to be preferred in all cases, as they will stand upright upon the table. Chemical flasks also have two patterns of lips, one, (fig. 11), simply turns outward, and this style is preferable either for pouring into or from the flask. The other lip is made by passing a heavy rim of glass around the top of the neck and afterward grinding the top evenly.

This style is preferable where corks are to be inserted. Let it be remembered that all chemical flasks should be made of *thin* glass,† and that in heating them the temperature at first must be gradually raised in the manner we have directed for evaporating dishes.

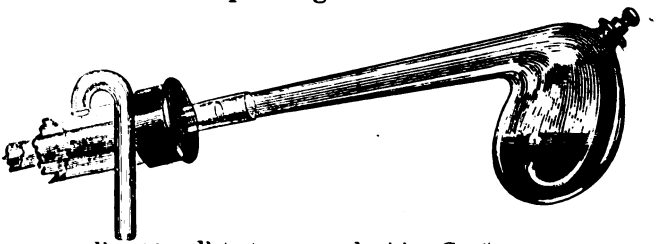

FIG. 12. Retort connected with a Condenser.

Retort.—Fig. 12 represents a retort connected with a condenser. This retort is known as the "tubulated and stoppered retort," and fig. 13 represents the "plain retort." It will be noticed that the distinguishing feature is the absence of a neck and stopper from the top of the plain retort. For all ordinary purposes tubulated and

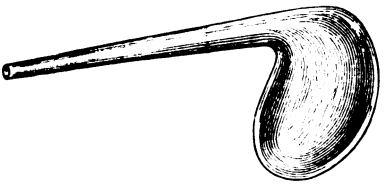

FIG. 13. Plain Retort.

stoppered retorts are to be preferred and the experimenter will seldom have use for the plain retort. In selecting a retort observe that the exit tube is regular and drawn rather small, so as to fit into the adapter without trouble.

†In selecting chemical ware to which heat is to be applied, whether the vessel be glass or porcelain, select that which is light and thin. The chemist is expected to be careful and not to break vessels by rough treatment. Contraction and expansion from variation of temperature is a matter beyond his control, and this will sooner or later fracture thick ware. It will also be found a matter of economy to purchase the best quality of glass and porcelain.

The body of the retort should be of thin glass and free from bubbles. Retorts are intended for the purpose of distilling, and should be warmed by means of a sand or a water bath. The liquid is to be introduced at the tubulure of the retort, and it is usually advisable to connect the retort with the condenser, and this with the receiver, before charging the retort.

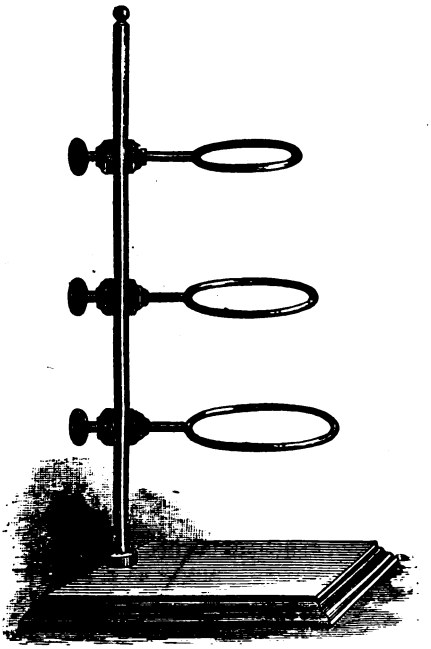

FIG. 14. Retort Stand.

Retort Stand.—Fig. 14 represents the common, ringed retort stand. It is arranged so that the rings may be moved up or down. The retort stand is used for holding a retort, or, indeed, any piece of apparatus, such as a flask, a sand-bath, a funnel while filtering, etc. In using the retort stand, the vessel to be held in position is placed upon one of the rings which is then moved into proper position and secured by means of the thumb-screw. Re-

tort stands are also made with clamps upon their arms by means of which test-tubes, glass tubing, etc., may be fixed in any desirable position.

Condenser.—A familiar example of the condenser is an ordinary worm, such as is used in distilleries. When heat is applied to the retort, the liquid changes into vapor, and passes into the condenser which is surrounded with cold water. The vapor consequently parts with its heat and is resolved into liquid, or in common language *condenses*, hence the name, "condenser." We present (fig. 15) a drawing of a condenser which is known as "Liebig's condenser." It consists of a tin or copper tube, two and one-half or three inches in diameter and two feet long,

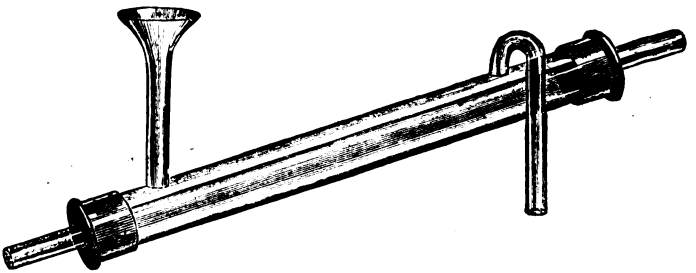

FIG. 15. Liebig Condenser.

corked at both ends. A glass tube passes through the center, extending beyond either cork. The outer tube (tin or copper), has attached near one end a half-inch metal tube; this inclines at an angle as represented by our engraving, and bears a funnel upon the top. At the other end of the large tin or copper tube another half-inch metal tube is affixed, and this passes around one side of the condenser and downward, as shown by our figure. Cold water is either poured into the funnel from time to time, or is permitted to flow into it in a steady stream. The water warmed by the heated vapor, consequently escapes at the exit. The retort is connected with the glass tube at the upper end of the condenser (see figure 12), and thus the tube through which the vapor passes is constantly surrounded with cold water.

Adapters.—These are shown by figures 16 and 17. They are designed to connect tubes of different sizes, to connect a retort with a condenser, or to connect the condenser with a receiver. It will be observed (fig. 16) that the straight adapter is simply a section of flaring glass tube. When tubes of different sizes are to be connected, the end of one is inserted into the large end of the adapter, and the other end of the adapter is to be inserted

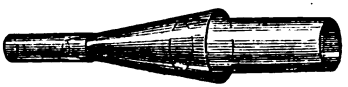

FIG. 16. Straight Adapter.

into the opposite tube, as illustrated by fig. 16. Curved adapters are bent as shown by fig. 17. They are used mostly to connect condensers with the receiver. A sheet of soft rubber should be wrapped closely around each joint and secured with string, as the adapter alone will not prevent the escape of vapor, or even of a liquid.

Receiver.—This is simply a bottle or a flask, unless the experimenter chooses to purchase the globular receiver. For ordinary purposes such is unnecessary, as a

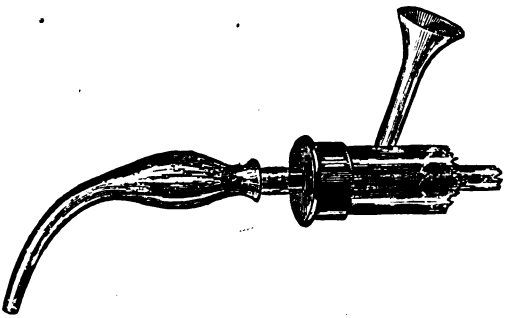

FIG. 17. Curved Adapter, connected with a Condenser.

condenser is used to liquefy the vapor of the distilling liquid. In this case the tip of a curved adapter inserted into an ordinary bottle will conduct the liquid to a place of safety. Refer again to fig. 17, which represents the

curved adapter attached to the exit of a condenser. If a bottle be placed beneath this, it will represent the exit of a distilling apparatus prepared for work, the bottle being the "receiver." In connecting an adapter with the receiver, the joint can scarcely ever be made air tight with safety. It is usually best to insert the tip of the adapter into the neck of the receiving bottle, and then to stuff cotton between the neck of the bottle and the adapter.

Glass Tubing.—This can be purchased of any size and weight. For all ordinary purposes it is best to select the weight which is employed in making homœopathic vials, or that which is a little heavier. The size of the tube must be determined by the work which the experimenter intends to perform, and the connections to be made. Glass tubes should always have their ends fused (after being cut) by means of the heat of an alcohol lamp or a gas flame. This is to prevent cracking, and to make it safe to handle them, and to insert them into corks, rubber tubes, etc.

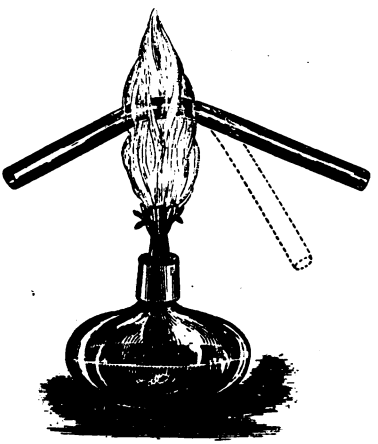

FIG. 18. Bending Glass Tubing.

Bending Glass Tubes.—Often it is necessary to bend glass tubing—a feat easily accomplished as follows: Hold the tube across the flame of an alcohol lamp until it gives

at the heated portion; then slowly slide the tube through the flame, bending it gradually as represented by fig. 18, the dotted line showing the intended bend of the tube. This will give a smooth, even curve, shown by fig. 19.

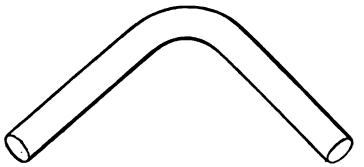

FIG. 19. Tube properly Bent.

If the flame be applied at one place, and the tube be then bent at this point, it flattens upon the outer side of the point of curvature, and the inner part of the curve often bulges upward. Such a bend will often almost completely close the bore of the tube, as we represent by fig. 20. The student will do well to experiment a little upon the bending of tubing.

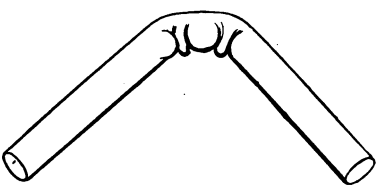

FIG. 20. Tube imperfectly Bent.

Cutting Glass Tubing.—Glass tubes may be cut at any place, by first filing a notch upon one side with an ordinary file, and then gently straining the tube between the fingers as though a stick were to be broken. The notch should be upon the side of the tube furthest from the operator, and against which the strain is directed. After the tube has snapped, burn each end in the lamp to smooth the edges.

Connecting Glass Tubes.—We have described the adapter. In addition to this chemists employ rubber tubing to connect glass tubes. For this purpose the soft rubber is used, and "black rubber" is more flexible than the white, and therefore to be preferred. If the rubber

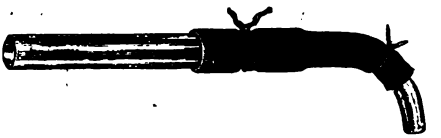

FIG. 21. Connecting Glass Tubes with Rubber Tubing.

is of larger bore than the diameter of the glass tube, it is necessary to fasten them by wrapping with string, as represented by fig. 21. If the bore of the rubber be the

FIG. 22.

size of the bore of the glass, simply inserting the glass into the rubber will make a secure joint (fig. 22). Formerly sheet rubber was used for making connections, but at present this is employed only in such places as forbid the use of rubber tubing. In almost all cases where glass tubes extend from one piece of apparatus to another, it is advisable to use two tubes, and connect them with rubber.

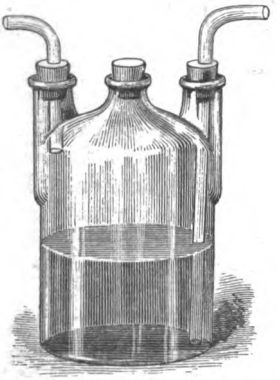

FIG. 23. Woulfe's Bottle.

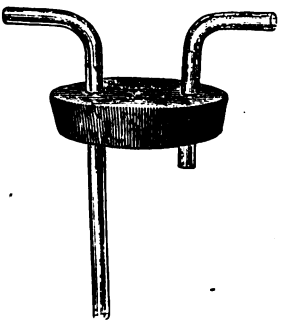

FIG. 24.

Woulfe's Bottle.—This is a bottle provided with either two or three necks. The Woulfe bottle is used for washing gases, or for holding a liquid while a gas passes into it. For this purpose one of the necks is fitted with a cork through which a glass tube extends and reaches nearly to the bottom of the liquid, and through this tube the gas is passed. If it be intended that this bottle be used simply for washing gases, the other neck is provided with a perforated cork, and through this cork a glass tube is inserted. In this manner several Woulfe's bottles may be connected together.

It is not necessary that the experimenter provide himself with Woulfe's bottles, as a wide-mouth bottle may be usually substituted in their stead. Through two holes in the cork the tubes may be passed, as we show by fig. 24.

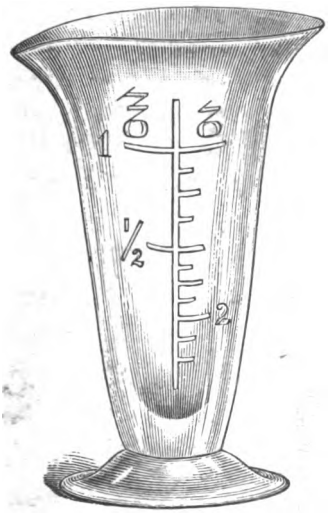

FIG. 25. Graduate.

Graduate.—The vessel to which the name "graduate" (fig. 25) is applied, is marked by a scale cut in the glass. This scale is graduated so that the operator may accurately measure either small or large amounts. For small

amounts the minim measure is employed. This holds one fluid drachm, and is graduated into fractions of this amount. The "graduate" proper will hold from half a fluid ounce to thirty-two fluid ounces, and is graduated from the drachm or half-drachm to that amount. Minims are marked *m*; drachms, ʒ; ounces, ℥.

Spatulas.—These are slender, knife-like steel blades, (fig. 26,) and are used for mixing powders, etc. There are three styles upon the market, metal handled, riveted wooden handled, and spatulas in which the blade is simply stuck into a wooden handle by means of cement. Always purchase the riveted wooden handled spatula.

FIG 26. Spatula.

Mortars are familiar to all. The ordinary wedgewood mortar answers best for ordinary purposes. They are not fitted for concussion, however, therefore an iron mortar must be used for breaking hard bodies.

other. The several substances have distinct properties, and this immutability man has never been able to overcome.†

Primary Forms of Matter, or the Elements.—All bodies that have never been decomposed into simpler forms of matter are called *primary bodies, or elements.* There are at present sixty-four known elements (see table).

Allotropy.—When this word is applied to an element we understand that the element is capable of existing in two or more distinct conditions. There is no change in composition, but the chemical and physical properties of the various modifications differ to a more or less extent. As an illustration we ask the reader to refer to phosphorus. This element may be made to assume three distinct forms, each of which differs from the others in a remarkable manner. The diamond, plumbago, and ordinary charcoal, are all carbon. Sulphur exists in several allotropic conditions, yet each form is sulphur. When an element assumes an allotropic form, we consider it simply a modification of the original body, regardless of the change in both chemical and physical properties. This change in the condition of an element is generally accepted as the result of a difference in the arrangement of its atoms.

Constitution of Primary Matter.—We have referred to the elements as though they were homogeneous bodies. We must now go further and assume that any perceptible amount of an element is capable of subdivision, until a given point is reached. This we consider the smallest particle of an element—*the atom*—and we accept that each form of primary matter is composed of atoms.

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† It must be remembered that we speak with the light which is now before us. At one time some of our common salts were called elements. The knowledge of a few years may prove that *our* elements are compounds. Indeed, the indications are that chemists may yet possess the power to decompose many of our so-called elements. However, if this should be so, and if all matter should be proven a modification of force, the chemist and theorist will be confronted with the problem as to the origin of *force*.

CHEMICAL THEORY, OR CHEMICAL PHILOSOPHY.

THE origin of the word *"chemistry"* is involved in obscurity. Some believe it to have reference to the soil of Egypt, and to denote something which is dark or black. Thus we hear the early science spoken of as *"the black art."* Other and equally good authorities believe that the word *chemistry* was derived in early days from a word which signified *heat*, and this also seems probable, inasmuch as the majority of chemical operations require the aid of heat.

Chemistry at the present day may be defined as that branch of natural science which treats of the properties of simple and of compound matter; of the relations of the various elements; of the compounds which may be formed of primary matters, and of the laws which influence combination (synthesis) and decomposition (analysis).

Matter.—Whatever has weight or occupies space is matter. Thus we may include in the same class hydrogen gas, the lightest, and osmium, the heaviest of known elements. All the material of this world is a form of matter, and its characteristic property is its indestructibility. Man admits that he can not destroy the smallest particle of matter, neither can he by any known power create. He has never noted the production of matter by an act of Nature, and Nature has never been known to annihilate. Man can bring no scientific evidence to disprove the assertion that all matter existed from the beginning, and will continue to exist forever.

Forms of Matter.—While it is admitted that matter is imperishable, we all know that many *forms* of matter exist. Thus we have iron, copper, gold, silver, lead, etc. These are primary forms of matter. They are classed under the general name, *matter*, but each differs from the

TABLE OF ELEMENTS,

WITH THEIR SYMBOLS, ATOMIC WEIGHTS, AND LARGEST QUANTIVALENCE.

	Symbol.	Atomic Weight.		Symbol	Atomic Weight.
Aluminium	Al^{vi}	27.3	Mercury, Hydrargyrum,	Hg''	199.8
Antimony, Stibium,	Sb'''	122.0	Nickel	Ni^{vi}	58.6
Arsenicum	As'''	74.9	Niobium	Nb^{v}	94.0
Barium	Ba''	136.8	Nitrogen	N^{v}	14.01
Beryllium	Be''	9 0	Osmium	Os^{iv}	198.6
Bismuth	Bi'''	210.0	Oxygen	O''	15.96
Boron	B'''	11.0	Palladium	Pd^{iv}	106.2
Bromine	Br'	79.75	Phosphorus	P^{v}	30.96
Cadmium	Cd''	111.6	Platinum	Pt^{iv}	196.7
Cæsium	Cs'	132.5	Potassium, Kalium,	K'	39.04
Calcium	Ca''	39.9			
Carbon	C^{iv}	11 97	Rhodium	Rh^{iv}	104.1
Chlorine	Cl'	35.37	Rubidium	Rb'	85.2
Cerium	Ce''	141.2	Ruthenium	Ru^{iv}	103.5
Chromium	Cr^{vi}	52.4	Selenium	Se^{vi}	78.0
Cobalt	Co''	58.6	Silver, Argentum,	Ag'	107.66
Copper, Cuprum,	Cu''	63.0	Silicon	Si^{v}	28.0
Didymium	D''	147.0	Sodium, Natrium,	Na'	22.99
Erbium	Eb''	169.0			
Fluorine	F'	19.1	Strontium	Sr''	87.2
Gallium	Ga	69.8	Sulphur	S^{vi}	31.98
Gold, Aurum,	Au'''	196.2	Tantalum	Ta^{v}	182.0
			Tellurium	Te^{vi}	128.0
Hydrogen	H'	1.0	Thallium	Tl'	203.6
Indium	In^{vi}	113.4	Thorium	Th''	231.5
Iodine	I'	126.53	Tin, Stannum,	Sn^{iv}	117.8
Iridium	Ir^{iv}	196.7			
Iron, Ferrum,	Fe^{vi}	55.9	Titanium	Ti^{iv}	48.0
			Tungsten, Wolfram,	W^{vi}	184.0
Lanthanum	La''	139.0			
Lead, Plumbum,	Pb^{iv}	206.4	Uranium	U^{vi}	240.0
			Vanadium	V^{v}	51.2
Lithium	Li'	7.01	Yttrium	Y''	93.0
Magnesium	Mg''	23.94	Zinc	Zn''	64.9
Manganese	Mn^{vi}	54.8	Zirconium	Zi^{iv}	90.0
Molybdenum	Mo^{vi}	95 6			

Atoms.—An atom is understood to be the smallest particle of an element capable of taking part in a chemical reaction. We have no means of judging of the size of atoms. We simply assume that such bodies exist and consider that each perceptible part of an element is made up of an innumerable number of atoms. Copper is supposed to be made up of atoms which have properties that distinguish the mass of copper from any other form of primary matter. This is true of gold, silver, oxygen, and the entire list of elementary bodies. We thus accept the old, old theory, that all matter is made up of atoms. We may to-day accept for the definition of the atoms, and perhaps not improve upon it, that made by Demokritos, 460 B. C., viz: "The atoms are invisible by reason of their smallness; indivisible by reason of their solidity; impenetrable and unalterable."

Atomic Weight.—In all works upon chemistry at the present day, we find after each element a number which expresses the atomic weight of that element, and which might be mistaken by the student for the *actual* weight of the atom of the element. The atomic weight of an element is accepted by chemists as the weight of an atom of that element, compared with an atom of hydrogen, but we have no means of determining the absolute weight of an atom of hydrogen, and we therefore do not know the actual weight of an atom of any element. As examples of atomic weight we give:

Antimony	122.
Uranium	240.
Copper	63.

The above indicates that an atom of antimony is 122 times as heavy as an atom of hydrogen. An atom of uranium is 240, and an atom of copper 63 times as heavy as an atom of hydrogen. Hydrogen is the lightest known body, the atom of hydrogen is accepted as unity (1). We may then define *atomic weight*, as the weight of the atom of any element, compared with an atom of hydrogen.

Combinations of Atoms (Molecules).—The atoms of elementary bodies are not believed to be in a free state. We consider that a bond (chemism) unites a certain num-

ber of atoms and this combination of the atoms of an element is called the *elementary molecule.* This molecule, then, is a combination of atoms which are alike.† The elementary molecule is supposed to be the smallest particle of an element that can exist in a free state.

Breaking up of the Molecule.—The atoms of an element are the indivisible constituents of its molecules. We have said that atoms do not (according to our present view) exist in a free state, excepting the few which are said to play the part of molecules. (See note below.) The breaking up, then, of the molecule of the element tends to mask the properties of the element. Example: The molecule of chlorine consists of two atoms of a kind, and the molecule of hydrogen also consists of two atoms thus:

HH represents the hydrogen molecule.
ClCl " " chlorine "

The above combination of atoms are elementary molecules, because they are formed in the one case by atoms of hydrogen, and in the other case by atoms of chlorine. Let us bring any even number of these molecules together in the sunlight and the positions of the atoms will be altered. An atom of chlorine unites with an atom of hydrogen, the result being two molecules in which the atoms are *unlike,* HCl. and HCl. Each of these molecules is a *compound molecule,* and the properties of the original bodies will have disappeared.‡

Compound Molecules, may be then defined as the smallest particle of a *compound,* and all compounds are composed of atoms that are unlike. We have, for example:

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†According to rather recent investigations the molecule of mercury and of nitrogen, each consists of *one* atom. Accepting this view we may say that the elementary molecule is the smallest particle of an element that can exist in a free state and perform the part of the element. This particle consists of one or more atoms, united in the latter case, by chemism.

‡ Some speak of compounds containing *elements.* Elements enter into compounds and may be again obtained from the compound. The element, however, loses its individuality when it passes into combination with another.

Chloride of Sodium.....NaCl.
Chloride of Potassium.....KCl.
Chloride of Hydrogen.....HCl.

and each molecule contains an atom of two different elements. We thus pass from the study of the primary form of matter to the study of:

Compound Matter.—The smallest particle of compound matter is a *molecule.* There can be no *atom* of any compound, for all compounds contain two or more kinds of primary matter, and therefore, the smallest conceivable portion of a compound must contain at least two atoms. As an example, we refer to bromide of potassium. Into any given amount of this compound the element bromine and the element potassium must enter. If we imagine that we have one hundred molecules of bromide of potassium, each molecule must contain one atom of bromine and one atom of potassium. We can separate the piece of potassium bromide into two parts, but each part will consist of molecules. We may theoretically subdivide the fragments into the one hundred molecules, and each molecule will have the properties of the original article, and further division would result in the production of atoms of bromine and atoms of potassium.

Primary Matter and Compound Matter Defined.

Primary Matter.	*Compound Matter.*
Atoms of one kind of primary matter unite to form elementary molecules.	Atoms of different kinds of primary matter unite to form compound molecules.
Primary matter consists of atoms that are alike, united by chemism.	Compound matter consists of atoms of two or more kinds of primary matter, united by chemism.
The smallest portion of primary matter is the atom.	The smallest portion of compound matter is the molecule.

Chemism.—There is an intimate connection between chemism, heat, electricity, and other forces, which is thoroughly discussed in works on physics. Chemism is a name which is given to the force which we believe produces the changes in the position of atoms. Chemical attraction, chemical affinity, and chemical force, are terms employed by chemists, denoting the same object, and all re-

fer to this force. Chemism is known by its effects. The constitutional transformation which is constantly taking place in matter, we credit to the action of this force. All forms of matter are subservient to it. The changes which it produces are infinite in number. Our very bodies grow and waste in obedience to this power, and yet the real force which we denominate *chemism* is unknown.

Effects of Chemism.—Chemical force destroys the distinctive properties of bodies, producing new substances. When elements unite, the resulting compound is different from either of the elements. For example, oxygen and hydrogen gases unite to form the liquid, water. If we heat sulphur and iron together, an entirely different substance results. In neither of the foregoing cases could we judge of the properties of the compound by the character of the elements which enter into it.

The antithesis of combination is decomposition; both depend upon the action of chemical force. Thus it is that from substances which are decomposed, bodies entirely different from the original may arise. Place, for example, a piece of sugar in a spoon, and hold it in the flame of an alcohol lamp. The sugar is decomposed, and the result will be a number of substances, none of which resemble sugar in the least. When a candle burns the tallow changes by combination with oxygen into (mostly) water and carbon dioxide, both of which are colorless, and differ from tallow and the oxygen of air in every way. Let it be remembered then that the result of chemical action, is altering of the properties of the substances concerned, with production of new bodies which have characters entirely distinct from those of the original.†

Mechanical Mixtures and Chemical Compounds.—When bodies are simply mixed, the properties of the component parts remain unimpaired. If for example we mix sand with peas, there is no alteration of the property of either substance. Each particle of sand re-

† It seems that elements which resemble each other may form bodies of somewhat similar character. Especially is this apparent in some alloys if we accept them to be distinct combinations.

mains intact and may be separated from the mixture in full possession of the original characteristics. If we mix powdered iron with powdered sulphur, the mixture will consist of particles of iron and particles of sulphur. The iron may be separated in its purity by means of a magnet, or the sulphur may be dissolved from the mixture with carbon disulphide. The foregoing are mechanical mixtures, and by the term we understand that bodies are simply in contact, with full possession of their natural properties. We take now the mixture of iron and sulphur, place it in a crucible and apply heat to it. Chemical action ensues and the particles of iron and of sulphur disappear. Neither the magnet nor carbon disulphide can extract or detect either of the original elements. In their stead we have the compound *sulphide of iron.* Thus we define and distinguish between substances which are simply mixed, and substances which are combined by chemism.

Equivalent Amounts.—The fact that elements unite to form compounds has been heretofore mentioned in general terms. The student might suppose that such combinations may result between any amount of the different elements. This is not the case. Combination ensues between definite quantities of the constituent elements. This fact has been long recognized, and thus the term *equivalent amounts or proportions* came to be used by chemists. The expression referred to the weight of one element of a compound as compared with the weight of another. Thus, it is found that 35.37 parts by weight of chlorine gas unites exactly with one part of hydrogen gas to form hydrogen chloride (hydrochloric acid). Therefore, the equivalent amount by weight of chlorine necessary to form hydrochloric acid by combination with one part of hydrogen gas is represented by the number 35.37. Again, it is found that 35.37 parts of chlorine by weight unite with 39.04 parts by weight of potassium. The equivalent amount, then, of potassium necessary to combine with 35.37 parts of chlorine is 39.04 parts. As an example of the practical application of the term equivalent we will say that the equivalent of iodine is 126.53, and the equiva-

lent of potassium is 39.04. Therefore, in making iodide of potassium, 126.53 ounces (or parts) of iodine must be employed for 39.04 ounces (or parts) of potassium. The foregoing is sometimes called *the law of constant proportions.*

Atomic Theory.†—This is based upon the theory of the atomic constitution of matter. According to this hypothesis chemical combination ensues between two or more atoms. One atom of hydrogen will unite with an atom of chlorine to form a molecule of hydrogen chloride. One atom of bromine will unite with one atom of potassium, the result being one molecule of bromide of potassium. The atoms of many elements unite in different proportions, thus: One atom of oxygen and two atoms of hydrogen unite to form a molecule of water. Two atoms of oxygen and two atoms of hydrogen unite to form a molecule of hydrogen peroxide.

Again: One atom of oxygen will unite with one atom of carbon to form one molecule of carbon monoxide. Two atoms of oxygen will unite with one atom of carbon to form one molecule of carbon dioxide, etc. According to the *atomic theory* all compounds, however complex they may be, are combinations of a definite number of atoms.

Molecular Weight.—The molecular weight of a body is the sum of the weight of all the atoms in the molecule. This we must accept when we acknowledge the correctness of our definition regarding the unalterable property of the atom and the indestructibility of matter. If 39.04 ounces of potassium are made to combine with 126.53 ounces of iodine, the resulting compound will weigh the

† Some speak of "the new Chemistry." The student should ever bear in mind that the act of chemical combination is to-day as it has been since man commenced the study of chemistry, and as we believe it always will be. Oxygen and hydrogen unite to form water in the manner they did when Lavoisier experimented upon the problem of the transformation of water into earth. There is no change in chemical facts. The term *new chemistry* simply refers to our theories. John Dalton resuscitated the atomic theory of the ancient Greek Philosophers and adapted the atom to his observations regarding chemical changes. This was the foundation for the work we now accept as the "New Chemistry."

sum of the two, viz: 165.57 ounces. If we imagine that the number 39.04 represents the weight of one *atom* of potassium, and that the number 126.53 represents the weight of one atom of iodine, the weight of one molecule of iodide of potassium must be 165.57. The molecular weight of a compound represents *the combined weights of all the atoms which enter into one molecule of that compound.* Thus, a molecule of chlorate of potassium will contain:

Potassium, one atom	39.04
Chlorine, one atom	35.37
Oxygen, three atoms, 15.96x3	47.88
	122 29

The molecular weight of chlorate of potassium, then, is 122.29, since this represents the combined weight of the five atoms which enter into its composition.

Combination by Volume.—We can not estimate the number of atoms in a given amount of any substance. The atom is a consideration beyond our senses as regards size and actual weight. In 1811 Avogadro, an Italian physicist advanced the hypothesis, *that equal volumes of any form of matter in the state of vapor, contain the same number of molecules* † (temperature and pressure equal).

While we can not estimate the number of atoms or of molecules in a given bulk of hydrogen gas, we can compare the weight of equal portions of different gases under like conditions, temperature, pressure, etc., with hydrogen. If we accept the hypothesis of Avogadro, we must accept that there should be the same relation between the molecular weights of the elements, that there is between the weights of different volumes of the elements while they are in the state of vapor. This is true, for accepting the weight of a given volume of hydrogen gas as 1, we find the weight of the same bulk of

† Avogadro made use of the term molecule, but in his day the distinction now drawn between the molecule and the atom was unknown. At that time the term atom and molecule were used indiscriminately to express theoretically the smallest particle of either simple or compound matter.

Nitrogen to be.....14.01
Oxygen to be.....15.96
Bromine to be.....79.75

and so on with the other gaseous elements. This corresponds exactly with the atomic weights of our table. It has been ascertained that whenever gases unite to form one substance the resulting compound in the form of gas or vapor measures *two volumes.* Thus:

One volume of chlorine and one volume of hydrogen form two volumes of hydrochloric acid.

One volume of oxygen and two volumes of hydrogen form two volumes of vapor of water.

One volume of nitrogen and three volumes of hydrogen form two volumes of ammonia.

Thus it is a rule that *whatever the volume of the gases before combination the result is always two volumes.*

Symbols.—Chemists employ either the first letter of the name of an element, or the first letter and another letter of the same word, as an abbreviation for the atom of the particular element. This is for convenience in writing and to facilitate the study of chemistry. Instead of writing the entire word carbon, the first letter, C, is used to denote that element, and whenever the student finds the letter C in a formula, he should understand that *carbon* is intended. The names of many elements commence with the same letter, thus: cadmium, cæsium, calcium, carbon, chlorine, cerium, chromium, cobalt, copper. To abbreviate these (excepting carbon) the first letter is associated with another letter of each word, and usually the Latin name is chosen, thus:

Cadmium.....Cd.	Cerium.....Ce.
Cæsium.....Cs.	Chromium.....Cr.
Calcium.....Ca.	Cobalt.....Co.
Carbon.....C.	Copper.....Cu.
Chlorine.....Cl.	

Of the above it will be noticed that the *u* of the symbol for copper is not present in the common name. The symbol Cu is derived from *cuprum,* which is the Latin

name for copper, and which does not agree with the English.†

Our table of the elements on page 27, gives, in the second column, the symbols of all the elements, and it will be noticed that there are several exceptions of that nature, as,

Antimony, Sb	from	*Stibium.*
Gold, Au	"	*Aurum.*
Mercury, Hg	"	*Hydrargyrum.*
Lead, Pb	"	*Plumbum.*
Silver, Ag	"	*Argentum.*
Sodium, Na	"	*Natrium.*
Potassium, K	"	*Kalium.*

We understand, furthermore, that each symbol represents *one atom*, therefore it is only necessary to see the symbol to recognize the element and the amount of the element. If we wish to indicate *two* atoms of potassium, we write a small figure 2 after the symbol, thus, K_2, three atoms would be expressed by K_3, four atoms, by K_4; etc. This is true of all the elements, for example:

One atom of oxygen	O.
Two atoms of chlorine	Cl_2.
Three atoms of bismuth	Bi_3.
Four atoms of carbon	C_4.

Some chemists prefer to place the figure at the upper right hand side of the symbol, or even, occasionally (when the symbol is alone in an equation) to place a large figure in front of the symbol, thus, C^2, or 2C, and these, in reality, amount to the same. In all cases a *single atom* is indicated by the symbol, and this is multiplied in accordance with the figure. The figure after the symbol refers only to that one symbol, and when several symbols are in a line the number of atoms of each separate element is represented. Thus, H_2SO_4 indicates, two atoms of hydrogen, one atom of sulphur and four atoms of oxygen.

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†We give the Latin names the preference in our table, making the English name secondary. This, we believe, will be acceptable to the majority of students. The author is of the opinion that the Latin names should be *the* names of our pharmacopœia, and indications point to the attainment of this object at some future day.

Chemical Formulæ.—A chemical formula represents the composition of a molecule, and is an association of the symbols of the elements which enter into the molecule. Thus, HCl impresses upon us the fact that one atom of hydrogen is united by chemism with one atom of chlorine, to form a molecule of hydrogen chloride (hydrochloric acid). H_2O is the formula for one molecule of water, and we see that in it two atoms of hydrogen are combined with one atom of oxygen. H_2SO_4 represents the composition of a molecule of sulphuric acid.

If we desire to represent more than one molecule, either a large figure is placed in front of the formula, or the formula is enclosed in parenthesis and a small figure placed after, thus:

$$2H_2SO_4 \text{ or } (H_2SO_4)_2$$

represents two molecules of sulphuric acid. In such cases the small figure after the parenthesis which incloses the formula, or the large figure in front of the formula, multiplies the entire formula or the atoms of each element in the formula. Examples:

$(BaO_2H_2)_4$ or $4BaO_2H_2$
$(KOH)_3$ or $3KOH$.
$(H_2O)_{10}$ or $10H_2O$.

Occasionally only a part of a formula is enclosed in the parenthesis. This multiplies the atoms which are represented by the enclosed symbols according to the foregoing rule, thus:

$Ba(OH)_2$ is the same as BaO_2H_2 or Ba2OH.
$Ca(OH)_2$ is the same as CaO_2H_2 or Ca2OH.

If we enclose a formula, part of which is in a parenthesis, in another parenthesis, part of the formula is multiplied in accordance with the figure after the outer parenthesis, and the part enclosed in the *two* parentheses is multiplied by the figures which follow both parentheses. Thus, $(Ba(OH)_2)_2$ represents two molecules of barium hydroxide: Barium, 2 atoms; oxygen, 4 atoms; hydrogen, 4 atoms. It is preferable, in expressing such formulæ, to place a large figure in front, as:

$2Ba(OH)_2$ instead of $(Ba(OH)_2)_2$.
$2Ca(OH)_2$ instead of $(Ca(OH)_2)_2$.

We understand, then, by the *formula* of a molecule, that the number of atoms of the different elements, as expressed by the symbols, enter into the composition of that molecule.†

Empirical Formulæ simply show us the elements which are in combination, with the amount of each. This will answer in many instances, but in actual work it is advisable that the formulæ of many molecules should impart an idea of certain characteristics of the molecule In making salts from organic acids, for example, it is best to show the number of atoms of basylous hydrogen. This is done by separating the hydrogen of the entire molecule, so that the number of atoms of basic hydrogen are separated from the hydrogen which is a part of the acidulous radical. When molecules are represented in this manner, instead of the term *empirical* formula, the term—

Primary Rational Formula is employed. As an example of the difference between the two modes of expression, we cite the following as illustrations:

	Empirical Formula.	Primary Rational Formula.
Tartaric acid	$H_6C_4O_6$	$H_2C_4H_4O_6$
Citric acid	$H_8C_6O_7$	$H_3C_6H_5O_7$
Malic acid	$H_6C_4O_5$	$H_3C_4H_3O_5$

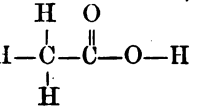
In all cases the number of atoms in the empirical formulæ and the primary rational formulæ is the same.

Graphic Formulæ are intended to represent the actual position each atom occupies in the molecule. Molecules are understood to have length, breadth and thickness. The graphic formulæ must always be expressed

† It must be remembered that the individuality of the elements disappears when compounds are formed. The formula of a molecule expresses the fact that definite amounts of certain elements entered into the molecule, and that by the destruction of this molecule the identical amounts of the same elements may be again obtained.

upon a plain surface, therefore graphic formulæ must be taken with this consideration. As an example of a graphic formula we give that of acetic acid.

Empirical formula of acetic acid, $C_2H_4O_2$.
Primary rational formula of acetic acid, $HC_2H_3O_2$.
Graphic formula of acetic acid,



Graphic formulæ are not used in this work.

Equations.—The equation is an arrangement of symbols to express any given chemical reaction. Equations are real chemical problems in which the signs +, —, and =, are employed exactly as they are in mathematics. If we wish to express the formation of iodide of potassium we write—

$$K_2+I_2=2KI.$$

Thus we see that one molecule (two atoms) of potassium unites with one molecule (two atoms) of iodine to form two molecules of iodide of potassium. The equation, $H_2+Cl_2=2HCl$, represents that when we add two atoms (one molecule) of hydrogen gas to two atoms (one molecule) of chlorine gas, after they combine two molecules of hydrogen chloride (hydrochloric acid) will exist. Sometimes equations are very complex, but all are formed upon the principle of those we have given.

The chemical equation is simply a symbolical expression to show the composition of bodies and the results which follow their decomposition.† In every equation

† Decomposition does not necessarily portend the separation of a compound into its elements. Thus, sulphuric acid is decomposed when we separate the characteristic acid radical (SO_4) from the hydrogen, which in this case acts as a base.

Double decomposition is a term often employed by chemists, and it denotes the re-arrangement of the atoms of two compounds, with the formation of other compounds. Thus when we mix solution of sodium carbonate with solution of calcium chloride, both salts are decomposed, and in their turn two new salts result (sodium chloride and calcium carbonate). Double decomposition often results in the formation of an insoluble substance, or one of difficult solubility in

there are always as many atoms after the sign of equality as there are before it. In the present work this should be the case in all instances, if the equation is correctly stated. Examples:—

$$\underset{\text{Sodium.}}{Na_2} + \underset{\text{Chlorine.}}{Cl_2} = \underset{\text{Sodium Chloride.}}{2NaCl}.$$

$$\underset{\text{Hydrochloric acid.}}{4HCl} + \underset{\text{Zinc.}}{Zn_2} = \underset{\text{Zinc chloride.}}{2ZnCl_2} + \underset{\text{Hydrogen}}{2H_2}.$$

$$\underset{\text{Tartaric acid}}{H_2C_4H_4O_6} + \underset{\text{Potassium hydroxide}}{2KOH} = \underset{\text{Potassium tartrate.}}{K_2C_4H_4O_6} + \underset{\text{Water.}}{2H_2O}.$$

The *parenthesis* is used with equations in the same manner as with formulæ. Thus, $(Zn+2HCl)_2$ is the same as $2(Zn+2HCl)$, and in both instances all the atoms enclosed are multiplied by 2.

Quantivalence, or Valence, is the property of an atom or power to hold in combination one or more atoms of other elements. Experiments show that all atoms do not possess the same combining power. An atom of hydrogen will combine with or take the place of an atom of chlorine, but it requires two atoms of chlorine for one atom of zinc. Hydrogen is taken as unity in this case, as heretofore, and is the standard.

If one atom combines with, or takes the place of

One atom of hydrogen, it is *univalent.*
Two atoms of hydrogen, it is *bivalent.*
Three atoms of hydrogen, it is *trivalent.*
Four atoms of hydrogen, it is *quadrivalent.*
Five atoms of hydrogen, it is *quinquivalent.*
Six atoms of hydrogen, it is *sexivalent.*

It is customary to speak of the element as a—

Monad element, if univalent.
Dyad element, if bivalent.
Triad element, if trivalent.
Tetrad element, if quadrivalent.
Pentad element, if quinquivalent.
Hexad element, if sexivalent.

---

the solution employed. The formation of salts from acids and alkalies, or acids and hydroxides, may be considered as examples of double decomposition; thus:—

$$\underset{\text{Sulphuric acid.}}{H_2SO_4} + \underset{\text{Potassium hydroxide.}}{2KOH} = \underset{\text{Potassium sulphate.}}{K_2SO_4} + \underset{\text{Water.}}{2H_2O}.$$

The quantivalence of an atom is noted in a chemical formula or an equation, by placing a dash at the upper right hand corner of the symbol, or by Roman numerals.† One dash represents a monad, two dashes a dyad, etc, as follows:—

Potassium is a monad,.....K′.
Calcium is a dyad,.....Ca″.
Boron is a triad,.....B‴.
Platinum is a tetrad,.....Pt^{iv}.
Niobium is a pentad,.....Nb^{v}.
Molybdenum is a hexad,.....Mo^{vi}.

It has been found that one atom of any monad element has the power to replace one atom of any other monad element in combination. Thus, one atom of potassium will displace one atom of hydrogen from the molecule of hydrochloric acid (HCl). One bivalent atom will generally displace two univalent atoms from compounds, or it will combine with two univalent atoms; for example:—

$$Ca'' + 2H'Cl' = Ca''Cl'_2 + H'_2.$$

One bivalent atom will generally displace or combine with one atom of any other dyad element, and so on.

In accordance with the foregoing, we find that—

Two univalent (′) atoms saturate one bivalent (″) atom.
Two bivalent (″) atoms saturate one quadrivalent (^{iv}) atom.
One trivalent (‴) atom saturates three univalent (′) atoms.
Two trivalent (‴) atoms saturate three bivalent (″) atoms.
One quadrivalent (^{iv}) atom saturates four univalent (′) atoms.
One quadrivalent (^{iv}) atom saturates two bivalent (″) atoms.
One quinquivalent (^{v}) atom saturates five univalent (′) atoms.
One sexivalent (^{vi}) atom saturates six univalent (′) atoms.
One sexivalent (^{vi}) atom saturates three bivalent (″) atoms.
One sexivalent (^{vi}) atom saturates two trivalent (‴) atoms.

Variation of Quantivalence.—The atoms of some elements have several combining powers. We can not offer an explanation of this fact. Sometimes the atom of

† The full quantivalence of each element, as we understand the matter at present, is given with each symbol of our table of elements (page 27). Atoms whose equivalence is odd are called *perissads*; atoms whose equivalence is even are called *artiads*. Thus, hydrogen, H′, gold, Au‴, and nitrogen, N^{v} are perissads; barium, Ba″, lead, Pb^{iv}, and nickel, Ni^{vi}, are artiads.

nitrogen is univalent, N′; again it is trivalent, N‴; and in other cases it has quinquivalent properties, N^{v}. This variation of valency always increases or diminishes by two, thus:—

Manganese.....Mn'', Mn^{iv}, Mn^{vi}.
Phosphorus.....P''', P^{v}.
Arsenic.....As''', As^{v}.
Nitrogen.....N', N''', N^{v}.

The student must be careful to distinguish between the marks which denote the quantivalence of an atom, and the Arabic numerals which designate the number of atoms in the molecule.

Analysis and Synthesis.—When two or more elements unite, they do so in exact and definite atomic proportions, and the properties of the original substances disappear. In their stead a new substance is found. The combination of elementary atoms to form compound molecules is *synthesis*. When compound molecules decompose, we have an *analytical* operation. Examples:—

1. Place a fragment of iodine in half a fluid ounce of warm distilled water, and add to the mixture a few iron filings. The iodine will disappear in a short time, and a colorless solution of iodide of iron (ferrous iodide) will be formed. This is a *synthetical* operation.

2. Place a small piece of bright copper in a solution of corrosive sublimate; the copper becomes coated at once with mercury and appears like silver. This is an *analytical* reaction for the corrosive sublimate is decomposed and mercury is determined.

We desire to call attention to the fact that the second example may also be classed with synthetical operations, for chloride of copper is formed as the mercuric chloride is decomposed. In actual work it is impossible to separate the two branches (analysis and synthesis). Chemists who prepare chemical compounds for commerce constantly decompose other compounds, and *analytical* chemists are continually forming new combinations.

Radical (root) chemical compounds are generally spoken of as though containing two parts, and each of these parts is known as a *radical*. This answers very well as

an aid to the study of chemistry, although such an arrangement may be only apparent. When one of the parts is an element, it is known as a *simple radical.* When a group of atoms takes the part of a single atom, we call the cluster a *compound radical.* Liebig defines the *radical* as follows:

"It is a never-varying constituent in a series of compounds.
"It can be replaced in these compounds by other simple bodies.
"In its compounds with a simple body this latter may be easily separated or replaced by equivalent quantities of other simple bodies."

Basylous Radicals.—These are usually metals, although hydrogen (see acids) and certain compounds of non-metallic elements act as such. To these the *acidulous* radicals attach themselves and the result is known as a salt. Examples of basylous radicals:—

Hydrogen..... H.
Potassium..... K.
Sodium..... Na.
Ammonium..... NH_4.
Calcium..... Ca.
Magnesium..... Mg.

Acidulous Radicals are usually non-metallic elements or combinations of non-metallic elements, although certain metallic oxides act as acids in presence of strong bases. Examples of acidulous radicals:—

Chlorine..... Cl.
Nitric acid radical..... NO_3.
Acetic acid radical..... $C_2H_3O_2$.

When acidulous radicals unite to saturation with basylous radicals, the properties of each disappear and a salt remains. Some bodies, as for example hydrogen, act as basylous radicals, and a part of compound acidulous radicals. In such cases it performs entirely different functions. Thus in H_2SO_4 it acts as a basylous radical, to which the acidulous radical (SO_4) is attached. In acetate of potassium ($KC_2H_3O_2$) hydrogen is a part of the acidulous radical, and in acetic acid ($HC_2H_3O_2$), hydrogen is the basylous radical, and a part of the acidulous radical. The combinations of atoms known as radicals, have a

The ordinary "acid cakes," a product of the manufacture of nitric acid from potassium or from sodium nitrate, are *acid salts* ($KHSO_4$ and $NaHSO_4$).

Haloid Salts.—Each of the *haloid* group (fluorine, chlorine, bromine and iodine,) of non-metallic elements will combine with hydrogen to form acids, thus:—

Hydrofluoric acid ..... HF.
Hydrochloric acid..... HCl.
Hydrobromic acid..... HBr.
Hydriodic acid..... HI.

When the hydrogen is displaced by a basylous radical, salts result, free from oxygen, and of which common salt (NaCl) is typical. This class of salts is sometimes distinguished as the *haloid salts.* Some authorities designate all of the salts which are free from oxygen as *haloid salts,* thus including such salts as potassium cyanide (KCN) and calcium sulphide (CaS). The term was originally derived from *sea salt.* The meaning (a likeness to sea salt) is rather favorable to the compounds of, or with, the chlorine group of elements. There is little use for the term at the present day.

Oxysalts or *oxyacid salts,* are terms once used, but now, deservedly, almost obsolete. Besides the haloid salts at one time a class of salts was recognized as being composed of an oxide of a metal and an acid. Thus:—

Sulphate of potassium was called the *sulphate of potash.*
Sulphate of magnesium was called *sulphate of magnesia.*
Ferrous sulphate was called the *sulphate of protoxide of iron,* or protosulphate of iron, etc. At this day there is no necessity for such a division. All of the metallic salts are considered as direct derivatives from the metals. The terms *haloid salts* and *oxysalts* should be dispensed with.

Compound Salts or *double salts,* are such as contain an acidulous radical in combination with two or more basylous radicals. We may also define double salts as those in which the basylous hydrogen of an acid is displaced by two or more basylous radicals (usually metals). Thus, $KNaSO_4$ is sulphate of sodium and potassium; $KNaC_4H_4O_6$ is tartrate of sodium and potassium, the ordinary "Rochelle salt."

Ate and Ite are terminations applied to salts which are made from acids which terminate respectively in *ic* and *ous*. Thus: Sulphur*ic* acid forms sulph*ates*; nitr*ic* acid forms nitr*ates*; sulphur*ous* acid forms sulph*ites*; nitr*ous* acid forms nitr*ites*.

There are exceptions to this rule, as for example:—

Hydrochlor*ic* acid gives rise to *chlorides*; hydrobrom*ic* acid give rise to *bromides*. The names of salts from these acids follow the rule that two elements united give rise to a salt terminating in the syllable *ide*. Thus:—

$$\underset{\text{Hydrochloric acid,}}{HCl} + \underset{\text{Potassium hydroxide.}}{KOH} = \underset{\text{Potassium chloride.}}{KCl} + \underset{\text{Water.}}{H_2O}.$$

$$\underset{\text{Hydrobromic acid.}}{HBr} + \underset{\text{Potassium hydroxide.}}{KOH} = \underset{\text{Potassium bromide}}{KBr} + \underset{\text{Water.}}{H_2O}.$$

It will be noticed that the bromide of potassium (KBr) and the chloride of potassium (KCl) are composed each of only two elements, although produced from acids which terminate in the syllable *ic*.

Ic and Ous are terminations which are applied to acids derived from the same element and containing severally more (*ic*) or less (*ous*) oxygen. Thus:—

Sulphur*ic* acid.....H_2SO_4
Sulphur*ous* acid.....H_2SO_3
Nitr*ic* acid.....HNO_3
Nitr*ous* acid.....HNO_2

At the present day the terminations *ic* and *ous* are used in a somewhat similar manner to distinguish classes of higher or lower oxides or salts, which are derived from a single metal. The class ending with *ic* contains more oxygen than the class ending in *ous*. Thus:—

Mercur*ic* oxide (higher oxide).....HgO.
Mercur*ous* oxide (lower oxide).....Hg_2O.
Ferr*ic* oxide (higher oxide).....Fe_2O_3.
Ferr*ous* oxide (lower oxide).....FeO.

In like manner the class of salts which corresponds with the higher oxide terminates in *ic*, and the class of salts which corresponds with the lower oxide terminates in *ous*. Thus:

Mercur*ic* chloride ($HgCl_2$,) contains more chlorine than mercur*ous* chloride (Hg_2Cl_2). Ferr*ic* chloride (Fe_2Cl_6) contains more chlorine than ferr*ous* chloride ($FeCl_2$).

Owing to the changes which have followed the study of chemistry, and to the many names which have from time to time been applied to compounds and to classes of compounds, a certain amount of confusion exists at the present day when we refer to the works of various writers.

Ide.—This termination indicates usually the combination of two elements, but occasionally denotes an element united to a radical such as *cyanogen*. Thus:—

Bromine and potassium form brom*ide* of potassium. Iodine and iron form iod*ide* of iron. Bromine and ammonium (NH_4) form brom*ide* of ammonium. Cyanogen (CN) and potassium form cyan*ide* of potassium.

The syllable *uret* was at one time used in place of the syllable *ide*, and occasionally is met with even now. As examples we name—sulph*uret* of potassium, now known as sulph*ide* of potassium. Cyan*uret* of potassium, now known as cyan*ide* of potassium.

Base.—"We may define a base as a *metallic oxide or hydrate capable of saturating acids.*" (Tilden.) The term base originally referred to the metallic oxide which was supposed to unite with the acid to form the oxysalt. The alkaloids are often called bases, but they differ from metallic oxides in their manner of forming salts with acids. They do not give rise to water as a product of the reaction.

Alkalies.—The name alkali, from *al* (the) and *kali* (a plant), was first applied to soda. At present, the hydroxides of all the potassium group of metals are called alkalies, and ammonia in addition thereto. Alkalies have the power of turning reddened litmus to blue, of neutralizing acids, and as a rule they are caustic in their action upon animal membranes. Ammonia is sometimes known as the *volatile alkali*, in contradistinction to caustic potash, soda, etc., which do not volatilize, and are called *fixed alkalies*.

Solution.—Water is the great solvent. When we speak of dissolving a substance, and do not name the liquid, water is always understood. We know very little about this phenomenon beyond the fact that certain solids, gases, and liquids will diffuse themselves to a fixed degree through liquids. For example, oxygen gas, sugar and sulphuric ether, will dissolve in water. Solution seems to differ from chemical combination, from the fact that the change is gradual and continuous, and not in definite proportions, as during chemical action. The liquid which dissolves the other substance is called the *solvent.* There is no known law regulating the act of solution, excepting the law of Dalton and Henry regarding pressure and the solubility of gases.† Some speak of dissolving zinc in dilute sulphuric acid, or of dissolving iron in hydrochloric acid, etc. Such expressions are improper, as the iron and the zinc do not dissolve, but form chemical compounds which are soluble.

Saturated Solutions.—A liquid is said to be saturated when it has taken up the solid liquid or gas to the fullest extent. The point of saturation is influenced by temperature. Most solids (a few exceptions, such as calcium citrate) dissolve to a greater extent in hot than in cold liquids. There is no law regarding the saturation of liquids. A liquid may dissolve one substance freely, and refuse to dissolve another of very similar chemical composition. Consequently the solubility of bodies in various liquids must be determined by direct experiment. The term *saturate* has a similar meaning when it is applied to the combination of an acid radical with a basylous radical or with an element. The term *neutralize* is used with the same meaning, and it is customary to speak of saturating (neutralizing) an acid with an alkali.

When we endeavor to more than saturate the bonds of

† "The weight of a gas dissolved by a liquid is directly proportionate to the pressure." (Law of Dalton and Henry.) Gases are more soluble in cold than in hot liquids (hydrogen excepted), and it is necessary to bear in mind that this is the reverse of the effect of heat upon the solution of most solids.

affinity, one substance is in excess, and remains unchanged. In this work the term "*slightly in excess*" will often be used. In such cases there is to be of one article a larger amount than is necessary to saturate the substance to which it is added. Thus: "Add diluted sulphuric acid to ammonia water until the acid is in slight excess," conveys the idea that the ammonia must be saturated with the acid, and the liquid afterward be of an acid reaction, from the effect of the free acid which is present.

Effervesce.—If a piece of marble or chalk, or a crystal of potassium bicarbonate, be covered with dilute hydrochloric acid, bubbles of gas (carbon dioxide) will escape, and the mixture is said to effervesce. If a bottle of carbonic acid water have its cork removed, effervescence at once takes place.

Fluorescence.—It is accepted that this phenomenon results from the change of highly refrangible dark rays of the solar spectrum into the less refrangible visible rays. If, for example, a few grains of sulphate of quinine be dissolved in a glass of slightly acidulated water, the solution will present a peculiar bluish cast. Some vegetable substances present this character to a wonderful degree. Gelseminic acid in minute proportion will render alkaline solutions fluorescent That product of the coal tar series which is known as "*uranine*" is one of, if not *the* most fluorescent of known bodies. One grain of this will render three hundred gallons of water distinctly fluorescent (greenish). The color produced by fluorescent bodies ranges from blue to green.

Effloresce.—When salts part with water of crystallization they are said to *effloresce*. Usually they fall into a dry powder under such circumstances. If a crystal of carbonate of sodium ($Na_2CO_3.10H_2O$) be exposed to dry air, it parts with nine molecules of water, a white powder, the mono-hydrate remaining ($Na_2CO_3H_2O$). Carbonate of sodium is thus said to be an *efflorescent* salt.

When a substance *absorbs* water from the air and liquefies, it is deliquescent (see *deliquescent*, page 80).

Isomerism.—When bodies contain the same elements, in the same proportion by weight, they are said to be *isomeric*. Thus, dextrine and starch are each $C_6H_{10}O_5$. Ammonium cyanate and urea are each CH_4N_2O. When isomeric bodies have the same constituents united so as to form molecules of different weight, the term *polymeric* is employed. Examples: acetylene, C_2H_2, and benzine, C_6H_6.

Heat.†—This is said to be the effect of molecular motion. Its general action is to expand simple matter, or even to destroy the attraction of its molecules. In this manner most simple bodies assume according to temperature a liquid or a gaseous state.

Heat is one of the most important of agents; it is indispensable to our existence, and is employed in most experiments in chemistry. By the proper application of this force the chemist is enabled to decompose most forms of compound matter, and to overcome the cohesive attraction of solids.

Effects of Heat.—Solids usually expand, liquefy, and then change into vapor, or pass at once from the solid to the state of vapor. Some substances, as for example iron, have an intermediate or pasty state; others, as lead, pass at once from the solid to the liquid. It seems that all forms of elementary matter might become solid if the temperature could be sufficiently reduced, and the pressure properly increased. Even the so-called permanent gases, as hydrogen and oxygen, may thus be considered vapors of solids, for recently both have been liquefied and frozen.

Latent Heat.—When a solid passes into the liquid form, a certain amount of heat disappears. This is popularly known as *latent heat*, and when the liquid again as-

† The doctrine of a fluid "caloric, or heat," once believed, is now discarded. Heat is no longer considered to be a material fluid, but a result of molecular motion. The study of the hypothesis as at present accepted belongs properly to works upon chemical philosophy and physics. In this work we shall necessarily refer to heat as though it were a tangible body, but the student is requested to bear in mind that such is not believed to be the case.

sumes the solid form, this heat becomes apparent. The amount of heat thus consumed in bodies which change from the solid to the liquid state varies with the element or the compound. The absorption of heat, or we may more properly say the heat consumed in overcoming the attraction of cohesion of the molecules of solids, may be illustrated by placing a thermometer in a vessel of mixed ice and water. If heat be applied, as long as the ice remains the temperature of the mixture can not rise above 0° C. (32° F.), and this is the original temperature of the mixture. As a rule, the change of state from solid to liquid is accompanied by absorption of heat. Freezing mixtures illustrate this, and a familiar example is a mixture of snow and salt, when, as both solids liquefy, a large amount of heat becomes latent.

Sources of Heat.—The great natural source of heat is the sun. Combustion always produces heat. Chemical combination, whereby a liquid assumes the solid form, is accompanied by the liberation of heat, as in slaking lime, in which case the water employed combines with the calcium monoxide. Intense chemical action results usually in the liberation of much heat, and often in combustion, as for example when a drop of sulphuric acid is placed upon a mixture of powdered sugar and powdered potassium chlorate. If sulphuric acid and water are mixed, heat results from the combination. When a body in motion is arrested, heat is evolved, thus we say that mass motion is changed into motion of molecules; example, strike a nail with a hammer, and the temperature of the nail is at once increased. Friction is a modification of the above, for force is continually being arrested by the moving substances which are in contact, and in proportion to the amount of friction the temperature of the surfaces rises.

Specific Heat.—In raising equal weights of different substances through the same number of degrees of temperature, different amounts of heat are required. Thus:—It requires 32 times the amount of heat to raise the temperature of one pound of water a certain number of de-

grees, that is required to raise a pound of mercury the same number of degrees. The "specific heat" of any body is the quantity of heat which is required to raise a pound of that substance one degree centigrade, water being taken as the unit. A very simple relation has been found to exist between the specific heat of elements and their atomic weights. This has given rise to the *Law of Dulong and Petit*, viz.: "The specific heats of the solid elements are inversely proportional to their atomic weights." This law then permits the deduction of the atomic weight of an element from a consideration of the *specific heat* of the element, and such calculations agree very nearly with the atomic weight as obtained by other means.

Measurement of Heat.—Thermometers are instruments designed to take the temperature of bodies. They are usually capillary glass tubes connected with bulbs of mercury. When the mercury expands it rises in the tube; when the mercury contracts, the column in the tube falls. We have said that heat expands solids and liquids, consequently the mercury of the bulb expands and contracts, as the temperature increases and diminishes. The measurement of heat by means of the ordinary thermometer is simply a register of the expansion of a liquid, as shown by the movement of the liquid in a capillary tube.

Centigrade Thermometer.—This is graduated so that the freezing point of water is zero (0°) and the boiling point 100°. This thermometer is accepted as authority by most chemists of Europe, but in this country the—

Fahrenheit Thermometer is usually employed by manufacturers, although scientists recognize the Centigrade. With the Fahrenheit scale the freezing point of water is 32° above zero (0°), and the boiling point of water is 212° above zero (0°). There certainly can be no excuse for this arbitrary scale, but such is the force of habit that it seems impossible to make general the adoption of the Centigrade scale in this country. In our work we gener-

ally give both scales; C. representing *Centigrade* and F. *Fahrenheit.*†

For chemical operations a cylindrical thermometer is used, the scale of which is enclosed in a glass tube. This can be employed for taking the temperature of liquids.

Temperature.—The temperature, accepted by chemists when none is mentioned, is 60° F. (15.5° C.) and that degree is understood unless otherwise noted. During the evaporation of liquids, and in other operations, the term "gentle heat" or "warmed," is often employed. This usually denotes a heat less than the boiling point of water, from 150° F. to 200° F. Where the term *evaporate* simply is used the heat of a water-bath is to be understood.

Evaporate until a pellicle forms, signifies that the evaporation is to be continued until a scaly substance or a skin forms upon the surface of the liquid. This may be illustrated by dissolving one part of bromide of potassium, or other salt, in water, then evaporating the solution until it is concentrated to saturation, when further evaporation results in the formation of a skin of the salt upon the surface of the liquid in the form of small imperfect crystals. This skin is the *pellicle.*

Light.—At present it is supposed that this agent is simply the undulatory movement of a medium which fills all space and pervades all matter. This medium is called ether, and it has been accepted as infinitely thin and elastic. Luminous bodies are supposed to cause this ether to vibrate, and these vibrations strike upon the eye and produce the sense of sight upon the same principle that waves in the air strike upon the ear and produce sound. The student should consult works upon optics and philosophy, as we can not enter into the details of the subject in this work.

Actinic Rays.—Besides the rays upon which burning bodies depend for their luminosity, known as luminous

†To convert Centigrade into Fahrenheit, multiply by 9, divide the product by 5, and add 32. To convert Fahrenheit into Centigrade, subtract 32, multiply by 5, and divide by 9.

rays, and besides the *heat* rays, a third radiant force exists, known as *actinic rays*, or *chemical rays*. These are invisible, and they accompany natural and artificial light rays. The art of photography and the changes which take place in certain forms of matter by exposure to light, result from their influence.

Spectrum Analysis.—All objects emit light if heated sufficiently. If a ray of ordinary light be passed through a prism it is refracted and dispersed, forming the colors, violet, indigo, blue, green, yellow, orange and red. These colors merge one into another and are not separated by dark intervals. If an incandescent gas be examined by the aid of an instrument prepared for the purpose, certain bright lines generally appear separated by dark bands. The position and color of these bands differ with the element or compound. An observance of the variations of this and connected facts has given rise to the means for detecting elements, known as *spectrum analysis*. The lines produced by the incandescent body is the *spectrum* and the instrument employed is the *spectroscope*. This instrument is carefully described by works upon chemical philosophy and physics, and by its use the constituents of the sun and stars have been determined. It is one of the most valuable and certain means for detecting small amounts of elementary matter, and the study of the spectroscope has given us several elements that otherwise could scarcely have been discovered.

Electricity.—When amber, glass, wax, shellac and certain other substances are rubbed with a dry cloth each body acquires a new property and has the power of attracting light objects, such as bits of cork, feathers, etc., and is said to be in a state of *electrical excitation, or electrified*. The force, which produces the attraction is known as the *electrical force*, or as *electricity*.†

It has been found that there are two states of electrical excitability, or as some choose to say, "two kinds of elec-

† It is customary for convenience to speak of the *electrical fluid*. This is improper, for the force which gives rise to the electrical state is unknown to man. We note the *effects* only of this agent.

tricity." That excited by rubbing glass is *positive* (or vitreous); that excited by rubbing wax, resin, etc., is *negative* (or resinous). The sign + always denotes positive electricity; the sign — denotes negative electricity. Bodies negatively electrified repel each other, and bodies positively electrified repel each other; but bodies in opposite electrical states attract each other.

Conductors.—Some bodies permit the ready passage of the electrical force; metals, water, etc., are examples, and are called *conductors*. Others, such as glass, resins, etc., are not capable of conducting electricity, or conduct it imperfectly, and are called *non-conductors*. Dry air is a non-conductor, air charged with vapor is a conductor. The ordinary electrical machine is simply a glass plate or cylinder, which is made to revolve against an amalgamated cushion and a strip of silk. The electricity excited is drawn from the opposite side of the revolving cylinder by means of brass points, and conducted to a Leyden jar. This form of electricity is often designated by the name *frictional electricity*, in contradistinction to—

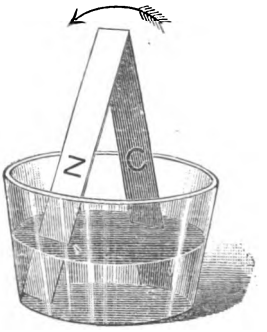


FIG. 27.†

† The chemical action at the surface of the zinc is thought to supply *negative* electricity to the zinc, and *positive* electricity to the acid. In practice but one current is recognized (positive) and when the two metals are connected at the top the "current" is described as passing from the copper to the zinc at the place of contact, and from the zinc to the copper through the solution (Fig. 27).

Voltaic Electricity, or Galvanism.— This was discovered by Galvani, and carefully considered by Volta, after both of whom it has been named. If a strip of zinc and a strip of copper be placed in a beaker glass partly filled with dilute sulphuric acid, and the two strips be connected with wire, or be brought in contact at the top, the zinc will dissolve, and bubbles will rise from the

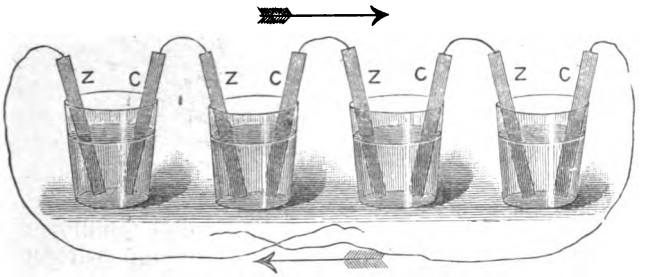


FIG. 28. Voltaic Battery.

copper. This chemical action is attended with the flow of an electrical current and the production in solution of zinc sulphate. If a number of such vessels be connected together by copper wire, the copper of one with the zinc of another (fig. 28), a *voltaic battery* is produced, and the electrical effect is intensified in accordance with the number of cells, for the current increases in strength with each additional cell. The opposite or end wires are known as *poles*, or *electrodes*; the wire which connects with the copper plate is the positive pole, while the term negative pole is applied to the wire which terminates the zinc.

There are many forms of the galvanic battery, all based upon similar considerations, but the metals and the solutions employed are variable and numerous, as are the forms of arrangement. A simple illustration of galvanic electricity may be shown by placing a silver coin beneath the tongue, and a strip of clean, bright zinc upon the tongue. If the edges of the metals are brought together over the tip of the tongue, a sharp metallic effect follows (some liken it to a burning taste), which ceases as soon as

the metals are separated. In this case the saliva of the mouth is the exciting liquid, and the sharp sensation results from the passage of the current of electricity through the tongue, which forms a part of the circuit.

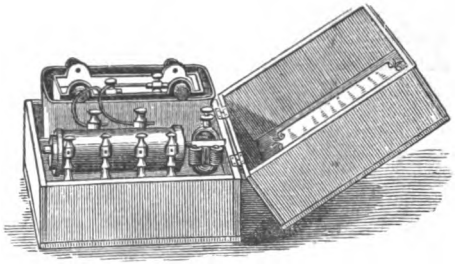


FIG. 29.

Galvanism at the present day is often employed by physicians, and many forms of the galvanic battery are manufactured expressly for their use. Fig. 29 represents the very complete and cheap battery designed by James Foster, jr., & Co., of Cincinnati.† The form of battery represented by fig. 30 is such as can be carried in the pocket; it is known as the Gaiffe battery.

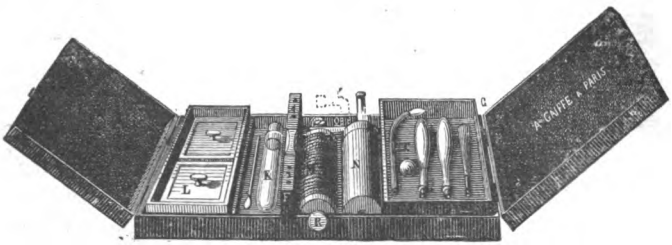


FIG. 30.

Electrolysis.—If the poles of a moderately strong galvanic battery are made of platinum wire and are placed near each other in a glass of dilute sulphuric acid,

† This is a form of the Faradic, or broken-current battery. We can not consume space with the principle of its action, and refer the reader to works devoted to electricity.

the current decomposes the water at the ends of the wires, and the operation is termed *electrolysis.* In this case it is found that pure oxygen escapes from the positive pole (electrode), and pure hydrogen escapes at the negative pole (electrode). If glass tubes are suspended over each pole, the two gases may be collected in the proportion of two volumes of hydrogen to one volume of oxygen. Compounds which are capable of such decomposition are called electrolytes.* The process (electrolysis) is applied to the analysis of certain salts in solution, or in the molten state. By this means several rare elements have been obtained, as will be stated in this work when elementary substances are individually considered.

Electro-Negative and Electro-Positive Elements.—In the electrolysis (analysis by means of electricity) of water, we have said that oxygen escapes at the positive pole, and hydrogen at the negative pole of the battery. Oxygen is thus the *electro-negative* element, and hydrogen the *electro-positive* element. As a rule, oxygen, chlorine, sulphur, etc., when combined with metals, appear at the positive pole, and are electro-negative. Hydrogen and the metals appear at the negative pole of the battery, and are known as electro-positive elements. These distinctions are to a certain extent relative, for under certain circumstances, dependent upon the elements combined, the order of liberation may be reversed. The following table of the electro-chemical elements (by Gore) is arranged so that "each substance is usually positive to all those following it, and negative to those preceding it."

---

* Those bodies only are electrolytes which are composed of a conductor and a non-conductor. Generally speaking, the greater the chemical opposition between the elements of a compound, the more readily it yields to electrolysis.—*Miller.*

Potassium,	Hydrogen,
Sodium,	Antimony,
Magnesium,	Carbon,
Zinc,	Tellurium,
Iron,	Phosphorus,
Aluminium,	Sellenium,
Lead,	Iodine,
Tin,	Bromine,
Bismuth,	Chlorine,
Copper,	Nitrogen,
Silver,	Sulphur,
Mercury,	Fluorine,
Platinum,	Oxygen.
Gold,	

We shall employ the terms electro-negative and electro-positive occasionally in this work, and in such cases the student is to understand that when the compounds named are submitted to electrolysis the electro-positive element is deposited, or escapes from the *negative* electrode or pole, while the electro-negative element is deposited upon or escapes from the *positive* electrode or pole.

Heat and Light by Means of Electricity.—The obstruction of a current of voltaic electricity results in the production of heat, or both heat and light. If the poles of a working battery be connected by a fine platinum wire the current of electricity is retarded, and in proportion to this obstruction heat and light are produced, the wire being heated to whiteness or even melted. If the poles of a strong galvanic battery be each connected with a carbon point, and after the circuit is completed the points be slightly separated, the result is a dazzling bright light. This is known as the *electric light*, or *electric arc*.

Magnetism.—Loadstone is a natural magnet. If a piece of loadstone be suspended by a fine thread one part will point to the north, and from whatever position it may be placed will settle into that direction. The portion which points toward the north is called the *north pole* of the magnet, and that toward the south is called the *south pole*. If a piece of steel be rubbed upon one pole of a loadstone the steel becomes magnetic. That part which was in contact with the pole of the magnet becomes the

opposite pole of the artificial magnet. The magnet thus produced will have poles the reverse of the magnet used to produce it. If the north poles of two magnetic needles be brought together a repelling force is exhibited and they fly from each other. This is true of the south poles of two needles. Upon the contrary the south pole of one needle *attracts* the north pole of the other, and if they are permitted to come in contact they adhere with considerable force. Thus we find that *like poles repel, and unlike poles attract.* In this work we occasionally refer to an element as capable of being magnetized, and in such cases it is to be understood that such a body will readily assume the character of a magnet. The force which gives rise to the property of polarity is unknown. We study its effects, sometimes speak of it as the *magnetic fluid*, but can not accept it to be a tangible body. It seems that most substances are more or less susceptible to the influence of magnetism.

Electro-Magnetism.—When an electric current is passed through an insulated wire wrapped about a bar of iron or steel, the bar becomes *magnetic.* When the current is cut off, if the bar be of soft iron, the magnetism immediately disappears; if the bar be of steel it remains a permanent magnet. This magnet has all the properties of the natural magnet and it is by this method that the ordinary "horse-shoe" magnets of commerce are prepared.

Specific Gravity.—The effects of gravitation are familiar to all. The force is unknown. All natural bodies have more or less weight, and it is customary to ascribe this property to gravitation. Specific gravity is "the weight of a definite volume of any solid, liquid, or gas, compared with the weight of a standard." We accept as the standard for solids and liquids water at the temperature of 60° F. (15.5° C.) If a given bulk of any body is heavier than the same bulk of water, we say its specific gravity is greater than that of water. If it be less than the weight of the same bulk of water, we say its specific gravity is less than that of water.

Specific Gravity of Solids Heavier than Water. Weigh the body in the air, then weigh it in water by suspending it by means of a hair from the arm of a balance, permitting the solid to enter the water. Divide the weight in air by the loss of weight in water, the result will be the specific gravity of the body under consideration.

The specific gravity of solids lighter than water is found as follows:—

Attach the light substance to a heavy body, such as a piece of brass, and find the difference of their combined weights in water and in air, which will equal the water displaced; subtract from this amount the weight of water displaced by the heavy body, (found by weighing it in air and then subtracting its weight in water,) and the result is the weight of water displaced by the lighter substance. Divide the weight of the light substance in air by the weight of water that it displaces, and the quotient is the specific gravity of that body. Where substances are soluble in water other liquids must be employed and the difference of their specific gravity noted.

Specific Gravity of Liquids—For this purpose a bottle holding exactly 1000 grains of distilled water, at the temperature of 60° F. is filled with the liquid under consideration and the weight of the liquid is taken. This is the specific gravity of the body; thus:—

If the bottle be filled with sulphuric ether U. S. P. this will only weigh 750 grains, and its specific gravity is 0.750, (water being 1000).

If it be filled with chloroform this will weigh 1490 grains, therefore, chloroform has the specific gravity 1.490.

If it be filled with mercury, the weight of this substance will be 13560 grains at 60° F. and consequently mercury has the specific gravity 13.5.

The *hydrometer* is a very convenient little instrument and is employed for taking the specific gravity of liquids where *exact* gravity is not required. Hydrometers are tubes of glass blown into a bulb at one end, into which a little mercury is introduced. When this instrument is placed in pure distilled water at 60° F. the heavy end sinks, the

tube remaining upright. The point at which the surface of the water comes into contact with the tube is marked 1 or 1.000, and is taken as unity. If now the tube be placed in a liquid of greater specific gravity than water, it will not settle so deeply; if it be placed in a liquid which is lighter than water it settles *below* the 1.000 mark. Hydrometers are provided with graduated scales within

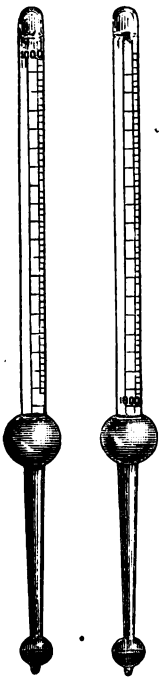


FIG. 31. Hydrometer for liquids heavier than water.

FIG. 32. Hydrometer for liquids lighter than water.

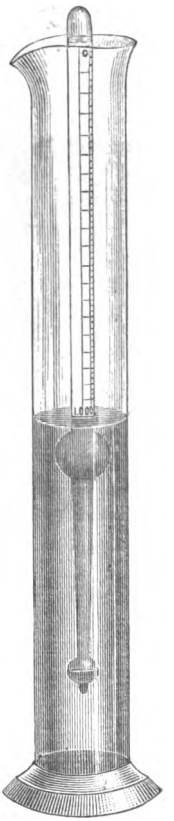


FIG. 33. Hydrometer tube with hydrometer.

the tube, and if the hydrometer settles to a given number upon this scale, that is the specific gravity of the liquid. This principle is derived from the fact that any body will displace its weight of a liquid, and if the liquid be dense (or heavy), the body floats higher than in a lighter liquid. Hydrometers for *heavy* liquids have the scale graduated from above downward, the 1.000 mark being at the top. Hydrometers for liquids lighter than water are graduated from below upward, the 1.000 mark being at the bottom of the scale. There is a narrow cylindrical glass made expressly for holding the liquid, which is known as the *hydrometer tube.*† (Fig's. 31, 32, 33.)

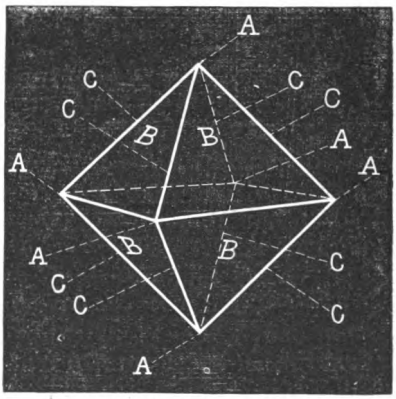


FIG. 34. B, faces of crystal; C, edges of crystal; A, angles of crystal.

Crystallography.—When a chemical compound passes from the liquid to the solid state it usually assumes a definite geometrical form. If for example a solution of common salt be evaporated to dryness it will be found that the residue is made up of a number of cubes or frag-

† Hydrometers are made and sold for use in the arts, under different names, such as *Saccharometer*, *Lactometer*, *Urinometer*, etc. Such are designed for particular liquids, and the scale is made with reference to each, but the general principle is that which we have named. We advise the use of hydrometers which recognize the true specific gravity of the liquid, instead of such gradations as are given by Baumé, Cartier, Twaddell, etc.

ments of the cube. Such a substance is said to *crystallize.* It is found that every chemical substance in the crystalline state possesses a distinct form.

Crystals are bounded by plane surfaces and these are called *faces.* Where the faces meet straight lines are formed and these are known as *edges.* The incidence of the faces forms *angles.* (Fig. 34.)

Crystals are bounded by either *similar* faces in which all the faces resemble and have the same relative position (figs. 35, 36), or they are bounded by *dissimilar* faces which occupy different positions and are of unequal size (fig. 37).

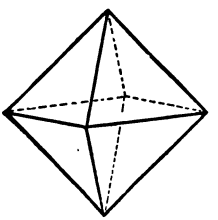
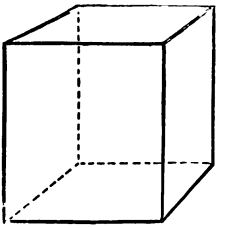


FIG. 35.
Regular Octohedron.



Similar Faces.
FIG. 36.
Cube.

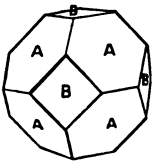


FIG. 37.
Dissimilar Faces.

A *simple* crystal is bounded by similar faces. A *complex* crystal is bounded by dissimilar faces, but the complex crystal is made up of two or more simple forms. *Dominant* faces of complex crystals are the largest, *subordinate* faces are the smallest. Thus A, A, A, A, of fig. 37, are dominant faces, and B, B, B, subordinate faces.

The majority of perfect crystals are of such a form that a face upon one side of the crystal is represented by a similar face on the other side. This, of course, is accompanied by a reproduction of opposite and similar angles and edges, so that a face, angle or edge, upon one side, is represented by a like face, angle or edge, upon the other side.

The *axes* of a crystal are imaginary lines drawn through the center of the crystal from center of dominant face to center of opposite face, or from angle to opposite angle. The particles of matter composing the crystal may be

supposed to be built up about these lines (axes), and thus the position and the length of the axes determine the position of the faces of the crystal.† (Fig's 38, 39.)

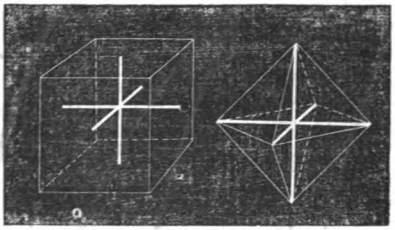


FIG. 38. FIG. 39.

Heavy white lines show the axes of the crystals.

Cleavage of Crystals.—Most crystals split more readily in certain directions than in others. This is termed *cleavage*. Ferrocyanide of potassium illustrates this fact most admirably, and the student is requested to examine a crystal of this substance. Our figure 40 shows the lines of cleavage whereby this salt may be split into thin slices.

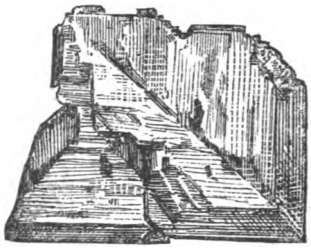


FIG. 40.

Systems of Crystallography.—There is an immense number of crystalline modifications, all of which are composed of combinations of six systems or groups. The forms of each system are closely related as regards

† From the fact that surfaces are *real*, and *axes* imaginary, it would seem natural to suppose that the better plan would be to assume that the *surfaces* and *angles* of a crystal determine the position and the length of the axes, and that the *axes* are imaginary skeletons upon which the faces are supposed to be constructed.

the length and position of their axes, and each system has a primary or *fundamental* form.

1. Regular or Cubic System.—This has three equal axes at right angles and of equal length. The fundamental form of this system is the regular octohedron and the next simple form is the cube (fig's 35, 36).

2. The Hexagonal System has four axes, three situated in one plane and of equal length, cutting each other at angles of 60°. The fourth axis is longer than the others, and intersects them at right angles to their planes (fig. 41).

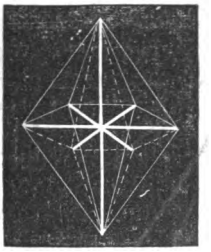


FIG. 41.

3. The Quadratic or Square Prismatic System.—These crystals have three axes, all of which are at right angles to each other, but only two are of the same length (fig. 42).

FIG. 42.

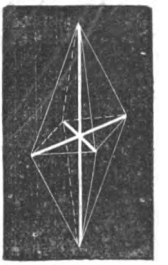


FIG. 43.

4. Rhombic System.—In this there are three axes of unequal length, all at right angles (fig. 43).

5. The Monoclinic System has three unequal axes, two of which are at right angles to each other, the third being inclined so as to be oblique to one of these and at right angles to the other (fig. 44).

6. The Triclinic System has three axes, all unequal in length and all oblique to each other. The crystals of this system are very irregular and difficult to study (fig. 45.)

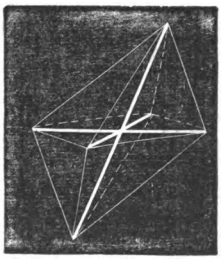


FIG. 44.

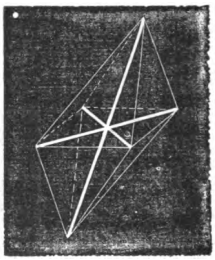


FIG. 45.

Determination of Crystalline Forms.—The form of a crystal is determined by means of an instrument named the *goniometer*. This gives the measurement of the angles formed by the faces, from which data the crystal is identified by simple calculation.

Permanent Crystals.—This term is often employed by chemists, in contradistinction to crystals which absorb moisture from the air and deliquesce, or which give off water of crystallization and effloresce. As examples of permanent crystals we may name bromide of potassium and chloride of sodium, either of which remains unchanged in ordinary air. Upon the other hand carbonate of sodium will *effloresce*, and acetate of potassium will *deliquesce*, therefore, these substances are not permanent. Some crystals undergo chemical decomposition, and such can not be called permanent.

Isomorphism.—When bodies have the same crystalline form, and agree in chemical composition to the extent that analogous elements or groups of elements can replace each other in such compounds, they are said to be *isomor-*

phous. Thus it has been found that many entirely different compounds have the same crystalline form, as for example the alums, and that they may be made to crystallize over each other without affecting the angles and faces of the crystal. In such cases the salts have, atom for atom, identical compositions. Examples:—

Potash Alum..... $KAl(SO_4)_2H_2O$.
Soda Alum..... $NaAl(SO_4)_2H_2O$.

Dimorphism and Trimorphism are terms applied to substances which have two or three distinct crystalline forms. Thus, if sulphur be allowed to crystallize from the molten state, it assumes the form of oblique rhombic prisms, while from solution in carbon disulphide, rhombic octohedra are formed. Many salts as well as many natural minerals have this property. If an element or a compound is capable of assuming *three* distinct crystalline forms it is said to be *trimorphous.*

Water of Crystallization.—Many crystals contain definite amounts of water in definite chemical proportions, and this water is known as *water* of *crystallization.* Water of crystallization often varies with the temperature at which the salt is crystallized, and the same salt may contain different amounts of water. Thus: Manganous sulphate ($MnSO_4$) crystallized at a temperature below 6° C. has the composition $MnSO_4+7H_2O$; between 7° and 20° C. it has the composition $MnSO_4+5H_2O$; between 20° and 30° C. a salt of the composition $MnSO_4+H_2O$ forms. Water of crystallization (saturated with the salt) often separates at ordinary temperatures, and during summer it is not uncommon to observe it running in quantities from barrels of sodium sulphate (Glauber salt). The temperature of 100° C. (212° F.) will expel the water of crystallization from most substances. (See effloresce.)

Experiments in Crystallization.—Dissolve nitrate of potassium in any given amount of hot distilled water until the salt ceases to dissolve. Place the solution in a cool place, and shortly crystals of potassium nitrate will appear, which will continue to increase in size until the

solution is depleted of the excess of the salt. This experiment may be performed with alum or bromide of potassium in the same manner and with the same result, excepting that each salt will assume a different crystalline form. After the solution has cooled it should be decanted from the crystals. The solution after crystallization of the salt is known as the *mother liquor*, a term often employed by chemists. If a saturated solution of most salts be permitted to evaporate spontaneously, crystals of the substance formerly in solution will remain. This is the principle applied in salt-making where the evaporation is performed by exposure to the sun's rays. Works upon chemistry usually state that it is necessary in order that large crystals be produced, to have the solution evaporate spontaneously. The author admits that large crystals may be made by slow evaporation, but he has made crystals of potassium bromide and other substances, of many ounces each in thirty-six hours, and many pounds in a few days.

Precipitates.—In chemistry the term *precipitate* is applied to an insoluble substance, which separates from a liquid, or which results when solutions are mixed. If we mix a little dilute ammonia-water with a little dilute solution of a ferric salt (ferric chloride, for example), a brownish substance will separate and settle to the bottom of the vessel. This is the *precipitate*. It does not necessarily follow that precipitates should settle (as the name signifies). Sometimes they rise to the surface, or even remain suspended in the liquid, in all cases being called by chemists precipitates.

Washing the Precipitate.—If the substance settles easily this operation is best performed by pouring off the supernatent liquid, filling the vessel with water, and after again permitting the precipitate to settle, pouring off the liquid as before. This operation may be repeated as often as is necessary to *wash* from the precipitate the substances held in solution by the original liquid, and the operation is known as "washing the precipitate by decantation." If the original mixture of precipitate and liquid be

poured upon a filter paper or a muslin strainer, the solution passes through and the insoluble matter (precipitate) remains within the paper or strainer. If water be now poured upon the precipitate the soluble matter will be carried through, and the precipitate will be purified. This is called "washing a precipitate upon a filter or a strainer."

Filtering.—There is a form of paper made expressly for filtering purposes. This is porous and when the object is simply to *filter* a solution, the operator is expected to pass the liquid through this paper. The act of filtration through paper simply separates insoluble or suspended matters from the solution, as these can not pass through the interstices of the paper. In some cases animal charcoal is employed for the purpose of filtration, especially with colored solutions. This results in the separation by absorption of some soluble constituent of the liquid.

Dialysis.—It has been found that crystallizable bodies in solution will pass through certain septa, as bladder and other animal membranes, gelatinous substances, etc., and that uncrystallizable bodies, as gums, can not do so. The application of this principle has led to the dialyses or separation of such as pass through the membrane (crystalloids) from those (colloids) incapable of doing so. This is known as *dialysis*. At the present day "dialyzed iron" is used in medicine to a considerable extent. This preparation is in reality a very basic solution of oxy-chloride of iron *incapable* of dialyzing and which remains inside the dialyzer, while the crystalline salts pass through. Dialysis has been applied to the separation of crystallizable poisons from organic mixtures.

Catalysis.—This term is sometimes employed by chemists in cases where a body induces chemical action and is itself unchanged. "Contact action" is occasionally used in the same manner. An illustration of catalytic action is that wherein finely divided platinum induces at ordinary temperatures the combination of oxygen and hydrogen, the platinum remaining unchanged. Some prefer to consider the liberation of oxygen from potassium

chlorate at a low temperature, if manganese dioxide be present, as an exhibition of catalytic action induced by the manganese dioxide. In this case the dioxide remains unchanged, all of the oxygen being liberated from the potassium chlorate. If the chlorate alone be used a very high heat is necessary to effect the liberation of oxygen.

INORGANIC CHEMISTRY.

NON-METALLIC ELEMENTS.

THE elements have been divided into two classes—*non-metallic* and *metallic*. This division is for the sake of convenience and is entirely arbitrary. A line can not be drawn between the two classes, and chemists are not agreed upon the position of some of the members, as for example arsenic, which is classed by many as a metal and by others as a non-metallic body. The term *metalloid* has been given to the non-metallic elements, but at present is seldom used. The non-metallic elements are usually electro-negative in combination with metals, therefore, most of their compounds are to be found in this work with the metals.

Non-Metallic Elements.—Hydrogen, oxygen, phosphorus, chlorine, sulphur, arsenic, bromine, selenium, boron, iodine, sellurium, carbon, fluorine, nitrogen, silicon. Chlorine, bromine, fluorine and iodine, are sometimes called *Halogens*. Each of the halogens can combine directly with bases, and the salts to which they give rise are often designated as *haloid* salts.

HYDROGEN.

Symbol, H. Atomic weight, 1.

Hydrogen is of peculiar interest to the chemist. It is the lightest known body and is taken as unity.† It was among the last of the so-called permanent gases, consti-

† It must be remembered that the *actual* weight of any element is unkown to us. We simply call hydrogen unity (1), and calculate the weight of other bodies by comparison with this accepted standard.

tutes a large share of all organic substances, both vegetable and animal, and forms about one-ninth part of the weight of the water of the Globe. Hydrogen seldom appears free, but a few volcanic emanations contain it in small proportion, and it has been found in meteoric iron, and in the gas of oil wells. Pure hydrogen may be obtained from water by electrolysis, but for ordinary purposes it is made by the action of dilute sulphuric acid upon metallic zinc. In this case it is contaminated with various foreign substances which give color to its flame, and an odor to the gas. Hydrogen may also be made by passing steam through a metal tube filled with iron turnings and heated to redness; black oxide (magnetic) of iron, known as ferroso-ferric oxide, Fe_3O_4, being also formed, thus:—

$$\underset{\text{Iron}}{3Fe_2} + \underset{\text{Water.}}{8H_2O} = \underset{\substack{\text{Black oxide}\\ \text{of iron.}}}{2Fe_3O_4} + \underset{\text{Hydrogen.}}{8H_2}$$

To make hydrogen conveniently the following experiment is readily performed. (Fig. 46). Into the 16 oz. chemical flask A, place two ounces of sheet zinc cut into small pieces. Cork the mouth of the flask with a cork fitted

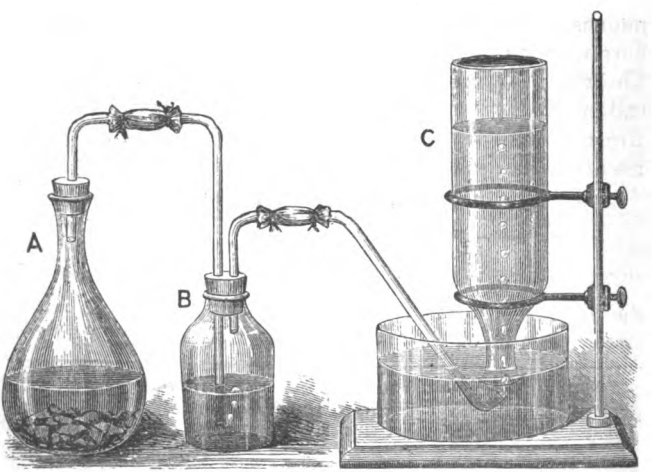


FIG. 46.

with a glass tube, the bore of which is not less than one-fourth of an inch, and which is connected with the 4 oz. bottle B. This bottle (B,) is to be half filled with distilled water, and the tube from the flask A. is to extend through the cork to near the bottom of the water. From the wash bottle B, a glass tube leads to the bottle C, filled with water, and which is inverted over a basin of water (pneumatic trough). Remove the stopper and pour into the flask A, a mixture of water seven fluid ounces, and sulphuric acid three fluid ounces. Replace the stopper and permit the gas to flow for a few minutes and escape from the exit of the tube, but not into the bottle C. Then move the terminus of the tube so that the flow of gas passes into the bottle (C) as indicated by our figure (46) until the water is displaced. The cork of the generator A, may be replaced by a cork with a straight glass tube, and after a few minutes from the commencement of the reaction the escaping gas may be ignited. Care *must* be taken that this gas is not lighted when the reaction begins, else a serious explosion may be expected. Hydrogen is colorless, tasteless and odorless, and was until 1877 classed among the *permanent gases*†. However, during the latter part of that year it was condensed by a pressure of 300 atmospheres accompanied by a low temperature. The apparatus employed need not be depicted here, but we take pleasure in referring the honor of this experiment to the simultaneous and independent work of M. Cailletet (France), and Raoul Pictet (Switzerland). The combining weight of hydrogen is less than that of any other body. Hydrogen will not support ordinary combustion, but it burns when a stream is lighted in the air or in oxygen, and the product is the vapor of water, thus:—

$$\underset{\text{Hydrogen}}{H_4} + \underset{\text{Oxygen}}{O_2} = \underset{\text{Water.}}{2H_2O}.$$

The heat produced by one gram of hydrogen combining with oxygen, as explained by the foregoing formula, is sufficient to raise the temperature of 34.462 grams of

† *Permanent gas* is a term which was once applied to gases which had never been liquefied. Oxygen and hydrogen are examples.

water from 0° C. (32° F.) to 1° C. (33.8° F). The following simple experiment will illustrate the production of hydrogen gas from water. Fill a test tube with water in a vessel of water, and raise it with the mouth down until the mouth of the tube is near the surface. Then place a pellet of sodium on the point of a penknife blade and quickly introduce it below the mouth of the tube. It will immediately fuse and rise into the tube and float upon the surface of the water, which will be decomposed (see sodium) with liberation of hydrogen. At last the tube will be filled with gas which will show the characteristics of hydrogen burning on the surface when a lighted match is thrust into it. Hydrogen will not support animal life when inhaled, but if the gas is perfectly pure it is said that a few inspirations may be drawn without danger. Hydrogen gas is sparingly dissolved by water and almost alike by water of all temperatures. This gas has the power to pass through many red-hot metals, such as platinum, palladium and iron. Oiled silk and india-rubber bags retain it but a short time. The absorption of hydrogen by hot metals has been termed *occlusion*. Many metals absorb and hold, at red-heat, large quantities of hydrogen, and such combinations have been likened to alloys. This form of hydrogen is by some supposed to be *solid*, and the term *hydrogenium* has been given it. The experiments of Troost and Hautefeuille, however, with palladium indicate that a real chemical compound is produced, and to such the composition PdH was ascribed. If hydrogen be mixed with oxygen gas or with atmospheric air and then ignited a violent explosion results. For this reason we have advised the experimenter to permit the gas to pass from the generator some little time before collecting it.

Experiment.—Thrust a lighted taper into a jar of hydrogen gas, the flame of the taper will be extinguished, but the surface of the hydrogen will burn with a pale blue light. Into an inverted tin tube, sealed at one end and pierced with a hole over which the thumb is held, pass hydrogen gas until the tube contains a mixture of air and hydrogen in about equal proportions; then cork the open end

and apply a flame to the orifice, a violent explosion results, (Fig. 47). In this instance the combustion is instantaneous throughout the tube, accompanied with formation of vapor of water. When a glass tube of one-half inch in diameter is held over the flame of a jet of hydrogen gas, a peculiar noise is produced. This, when in operation, is called the *musical tube.*

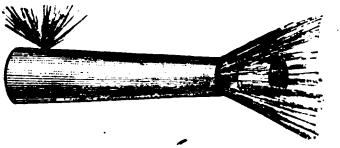


FIG. 47.

Hydrogen gas is used in medicine only in combination, but in this form it is of vital importance to both animal and vegetable life. The *oxy-hydrogen* blow-pipe flame is dependent on combustion of mixed jets of oxygen and hydrogen gases, and is of great importance in the arts and in chemical analysis as a source of intense heat. Hydrogen gas is sometimes employed for filling balloons, but for economy's sake coal gas is usually used. Hydrogen combines with elements to form *hydrides.* The acids are salts of *hydrogen.*

Hydrogen Monoxide (Water).—Formula, H_2O. Molecular weight, 17.96.

The molecule of water is composed of two atoms of hydrogen united by chemism with one atom of oxygen, or, two volumes of hydrogen in combination with one volume of oxygen. We have shown that water may be decomposed by sodium, yielding hydrogen gas as one of the products. If it be decomposed by electrolysis, two volumes of hydrogen gas rise from the negative pole for one volume of oxygen from the positive pole, and it will be found that the weight of the oxygen as compared with the hydrogen is as oxygen 16, hydrogen 2. The bulk, however, of the hydrogen, and the number of hydrogen atoms, are each *twice* that of the oxygen.

Pure water is seldom to be procured in our large cities, even by distillation. Most natural waters contain organic matter, either volatile or which decomposes by the heat of distillation, and thus contaminates the distillate with certain products which distill and condense with the steam. The air of cities and of dwellings, more especially of chemical laboratories and of manufactories, swarms with organic matter, germ cells, etc. If, during the act of distillation in these places, the condensed water is permitted to fall through the air, even from the condenser to a receiver, the result is a liquid which contains all that is necessary for the propagation of the lower classes of organic vegetable life. This fact is exemplified by the "slimey" contents of our pharmacist's bottles of distilled water, after a little time. Distilled water (not absolutely pure, but such as will answer all ordinary purposes), should be made by connecting a retort with a Liebig condenser, the exit of which connects with a bent glass tube which enters into a bottle through a plug of cotton placed in the neck. The retort should be two-thirds filled with fresh spring water, of which one-fifth is distilled and thrown away. The receiving bottle should now be connected with the condenser, and distillation conducted until three-fifths have passed. The distillation is now to be discontinued and the receiver securely stopped. The water of many mountainous lakes is very pure, especially as regards organic matter, and so we may say of some Northern springs.

Water is tasteless and odorless. It is colorless if in small body, but if in large quantities is bluish-green. This is particularly noticeable in the waters of certain pure lakes which are fed by springs or mountainous streams. Water is almost incompressible and is a bad conductor of heat and electricity. It is remarkable, as being an exception to the general rule that bodies contract upon cooling and expand when heated, for when water is heated from 0° C. to 4° C. (32° F. to 39.2° F.) it *contracts*, and when it is cooled from 4° C. to 0° C. (39.2° F. to 32° F.) it *expands*. This is a fact of the utmost importance to man, inasmuch as upon this principle ice forms upon the

surface of water and remains there instead of sinking. In the act of cooling, from the surface, water constantly sinks and is replaced by warm water from below, until the entire body is reduced to the temperature of 4° C. (39.2° F.) Then it congeals on its surface, and the loss of heat from that time must be by *conduction*, and, as we have said, water is a poor conductor. Therefore it is that our lakes and rivers seldom freeze to a very great depth.† The freezing point of water is usually 0° C. (32° F.) but under certain conditions it may be reduced as low as —24° C. (11.2° F.) without solidifying. Snow and frost are forms of ice crystals.

Water boils at 100° C. (212° F.) at the level of the sea. As the pressure of the atmosphere is lessened (as in ascending a mountain) the degree at which it boils becomes proportionately less. Water is the universal solvent. In all cases in this work where the direction is to make a solution, no liquid being designated, water is understood.

Water combines chemically with many basic oxides, to form a class of compounds termed *hydroxides*, *hydrates*, or *hydrated oxides*. In these the elements of water (H_2O) are present, but they do not exist as water. Examples:—

$$\underset{\text{Potassium monoxide}}{K_2O} + \underset{\text{Water.}}{H_2O} = \underset{\text{Potassium hydroxide.}}{2KOH}.$$

$$\underset{\text{Calcium monoxide}}{CaO} + \underset{\text{Water.}}{H_2O} = \underset{\text{Calcium hydroxide.}}{Ca(OH)_2}.$$

† In Northern countries a kind of ice forms in *exceedingly* cold weather upon the rocks upon the bottom of rapid ripples, while running water is above and *no* surface ice. This ice is known under the name of *anchor ice*. It, in appearance, resembles masses of needles with their points up, and attached at their bases to the rocks. We can account for this ice by the fact that the mass of the water of the ripple is almost at the freezing point. The rocks conduct the heat of the bottom of the ripple to the bank and thus actually reduce the temperature of the rocky bottom to less than 0° C (32° F.) This is followed by congelation upon the bottom in crystals which shoot upward into the stream in the manner we observe crystals shoot across the surface of water in the roadside puddles. It is well known in the North, that *anchor ice* portends very cold weather, and it is customary to say, in order to make the impression of severely low temperature, *anchor ice is forming*. Anchor ice is lighter than water and floats when it is broken from the rocks.

Some of the hydroxides (hydrates) hold the elements of water with great force, a very high heat being required to dissociate them.

Hygroscopic substances are those which absorb moisture from the atmosphere, and retain it as water. This is true to an extent with most dry powders in a moist atmosphere, but the term *hygroscopic* is applied in this work to such bodies only as have a positive attraction for moisture in ordinary air.

Deliquescent.—If a body absorb moisture in sufficient amount to dissolve in the water so taken from the air, it is said to *deliquesce*. All deliquescent bodies are hygroscopic, but it does not follow that a hygroscopic substance is deliquescent. As an example of deliquescent substances, we can name acetate of potassium, or chloride of calcium, either of which will absorb moisture from the air upon exposure to it, and dissolving in this water will form a liquid.

Anhydrid.—Oxides which will unite with water to form acids are called *anhydrides*. Thus:—

$$\underset{\text{Sulphurous anhydrid.}}{SO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Sulphurous acid.}}{H_2SO_3}.$$

$$\underset{\text{Nitric anhydrid.}}{N_2O_5} + \underset{\text{Water.}}{H_2O} = \underset{\text{Nitric acid.}}{2HNO_3}.$$

$$\underset{\text{Carbon anhydrid.}}{CO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Carbonic acid.}}{H_2CO_3}.$$

The anhydrid of sulphuric acid is in reality SO_2. This will unite with hydrogen dioxide (H_2O_2) to form sulphuric acid, thus:—

$$\underset{\text{Sulphur anhydrid.}}{SO_3} + \underset{\text{Hydrogen dioxide.}}{H_2O_2} = \underset{\text{Sulphuric acid.}}{H_2SO_4}.$$

Owing to the fact that the real anhydrid of sulphuric acid can not form the acid by combination with water, it is customary to speak of sulphur trioxide (SO_3) as the anhydrid, thus:—

$$\underset{\text{Sulphuric anhydrid.}}{SO_3} + \underset{\text{Water.}}{H_2O} = \underset{\text{Sulphuric acid.}}{H_2SO_4}$$

Anhydrous bodies are those which do not contain water. Salts are *hydrous* when they contain water that may be driven off without altering the chemical properties of the compound to any considerable extent.

Hydroxyl, OH. This is the foundation of all hydroxides. We may consider the molecule of water H_2O to consist of hydroxyl (OH) united to hydrogen forming *hydroxide* of *hydrogen*, or water, H.OH. The atom of basic hydrogen may be displaced by other univalent atoms, and thus from the molecule of water the combination OH may be carried and serve as the foundation for a complicated series of compounds. The formation of the hydroxides from hydroxyl may be illustrated as follows:—

$$\underset{\text{Hydroxyl.}}{2OH} + \underset{\text{Hydrogen.}}{H_2} = \underset{\text{Hydroxide of hydrogen (water.)}}{2HOH}.$$

$$\underset{\text{Water.}}{2HOH} + \underset{\text{Sodium.}}{Na_2} = \underset{\text{Hydroxide of sodium.}}{2NaOH} + \underset{\text{Hydrogen.}}{H_2},$$

The term *free hydroxyl* has been applied to hydrogen dioxide (peroxide of hydrogen) H_2O_2. Such a term, however, is not accepted by chemical writers. Compounds like calcium hydroxide, $Ca(OH)_2$, are considered as two univalent molecules of OH (hydroxyl), in combination with one bivalent atom.

Hydrogen Dioxide, or Peroxide of Hydrogen.

Formula, H_2O_2.

This was discovered by Thenard in 1818. It may be made by acting upon barium dioxide (peroxide of barium) with hydrochloric acid. It is a colorless, caustic liquid, soluble in ether, and the ethereal solution may be distilled without decomposition. Hydrogen dioxide is of no practical importance in medicine, but it has been used as a bleaching agent, and is said to be the liquid which is employed by some women to change the color of their hair to yellow (blonde).

ACIDS—Salts of Hydrogen.

In 1816, Dulong propounded the theory that all acids are combinations of hydrogen and an element or a group of elements. This is now the accepted view of the majority of chemists, consequently we shall consider acids in this place, as they properly belong to the hydrogen compounds. From necessity we introduce the organic acids which are used in medicine, as well as the inorganic acids.

Acids.—Acids are a class of compounds which are generally sour to the taste, hence the name *acid*, a word derived from a Latin root denoting an edge. At one time it was supposed that *all* acids were sour, reddened litmus and certain other vegetable blues, and neutralized alkalies. Lavoisier considered all acids to be compounds of oxygen, which idea has since been refuted, although at the present day only about six acids are known that do not contain oxygen. The element common to all known acids is *hydrogen*, and at this time it is customary to speak of acids as *salts of hydrogen;* thus :—acetic acid is known as acetate of hydrogen, sulphuric acid as sulphate of hydrogen, etc. This definition of an acid seems reasonable as we now view the science of chemistry, for each acid contains a fixed amount of hydrogen which is capable of being replaced by a metal. Thus, sulphuric acid (sulphate of hydrogen) H_2SO_4 is decomposed by zinc, Zn, to form sulphate of zinc, $ZnSO_4$, the hydrogen H_2 being liberated. The radicals of acid molecules are bodies linked by chemism with the hydrogen, which (hydrogen) is exchangable for metals, (page 43). Sometimes the acidulous radical is an element, thus:

Hydrochloric acid...(Chloride of hydrogen)...HCl,
Hydrobromic acid...(Bromide of hydrogen)...HBr,
Hydriodic acid.. .....(Iodide of hydrogen).....HI,

are acids, in each of which the hydrogen is united with one atom of the respective element, which element, in this case, is the *acidulous radical.* Usually, however, the acidulous radical is a combination of elements which have a strong affinity among themselves, but are loosely attached to the basylous hydrogen, thus :—

Sulphuric acid, H_2SO_4, and Carbonic acid, H_2CO_3, contain respectively in each molecule two atoms of hydrogen capable of being replaced by metals, the remainder of the molecule constituting the *acidulous radical;* consequently SO_4 is the acidulous radical of sulphuric acid; and CO_3 of carbonic acid.

Basylous Hydrogen.—In the foregoing illustrations the entire amount of hydrogen in each molecule is replaceable by metals. This hydrogen acts the part of a base to which the acidulous radicals are attached. Many of the organic acids, however, contain hydrogen as parts of the *acidulous* radical. We may mention—

Citric acid ..... $H_3C_6H_5O_7$,
Tartaric acid..... $H_2C_4H_4O_6$,
Acetic acid..... $HC_2H_3O_2$,

and in all such the hydrogen entering into the acidulous radical can not be removed without destroying the molecule, and the properties of the compound. Thus, replace one atom of hydrogen of the molecule of acetic acid with an atom of potassium, and we have a molecule of acetate of potassium $KC_2H_3O_2$. Remove now by means of heat another hydrogen atom or atoms and the characteristics of the compound are utterly destroyed. From acetate of potassium we may re-obtain the acetic acid radical, but from the result of the decomposition by heat it is impossible, for this acidulous radical no longer exists. We may thus define basylous hydrogen in these instances as that portion of the hydrogen in the molecules of acids which is capable of being replaced by a metal without destroying the properties of the acidulous radical. (See radicals, page 43).

Quantivalence of Acids.—Some acids contain one atom of hydrogen capable of being replaced by metals, others two, three, etc. Thus, hydrochloric acid (HCl), and acetic acid ($HC_2H_3O_2$), each contains *one* replaceable atom of hydrogen in each molecule, and is *univalent;* sulphuric acid (H_2SO_4), and tartaric acid ($H_2C_4H_4O_6$), liberate *two* atoms of hydrogen, and are *bivalent;* citric acid

($H_3C_6H_5O_7$) contains in each molecule, *three* atoms of hydrogen that may be displaced without destroying the properties of the acidulous radical and is hence *trivalent*. The terms mono-basic, dibasic and tribasic, are applied to acids in accordance with the number of atoms of basylous hydrogen, thus:—

Formic acid.....$HCHO_2$ .....is monobasic.
Tartaric acid.....$H_2C_4H_4O_6$.....is dibasic.
Citric acid.....$H_3C_6H_5O_7$.....is tribasic.

Formulæ of Acids—We have shown that acids may with propriety be called salts of hydrogen.† In view of this fact we prefer, when possible, to express their composition by means of formulæ in such a manner as to show their quantivalence. Thus, the actual composition of formic acid, (empirical formula), is H_2CO_2; but experiment has demonstrated the molecule of the formic acid radical to be *univalent*; therefore, chemists separate the atoms of hydrogen and represent the molecule $HCHO_2$. This latter mode of expression is called a *primary rational formula*. (See page 38). In illustration of this we call attention to the formulæ of the following acids, a glance at the primary rational formulæ of which will show the acidulous radical united with the hydrogen that may be displaced by metals, while the *empirical* formulæ as expressed by the second column is comparatively useless:

	Primary Rational Formulæ.	Empirical Formulæ.
Acetic acid,	$HC_2H_3O_2$.	$H_4C_2O_2$.
Tartaric acid,	$H_2C_4H_4O_6$.	$H_6C_4O_6$.
Citric acid,	$H_3C_6H_5O_7$.	$H_8C_6O_7$.
Malic acid,	$H_3C_4H_3O_5$.	$H_6C_4O_5$.

Medicinal Acids.—There are many hundred known bodies classed among the acids, but of these 42 only are to any extent used in medicine either in substance or in combination, unless it be as parts of plants or fats.

† Dulong in 1816 was the first to make this definition, although Davy and Gay-Lussac had previously shown that hydrochloric acid and hydriodic acid are free from oxygen.

ACIDS OF INTEREST IN MEDICINE.

Acetic,	Malic,
Arsenic,	Nitric,
Arsenious,	Nitro-muriatic,
Benzoic,	Nitrous,
Boric (Boracic),	Oleic,
Carbazotic,	Oxalic,
Carbolic,	Phosphoric,
Carbonic,	Meta-phosphoric,
Chloric,	Phosphorous,
Chromic,	Pyrogallic,
Chrysophanic,	Pyroligneous (mixture),
Citric,	Pyrophosphoric,
Formic,	Salicylic,
Gallic,	Silicic,
Hydriodic,	Succinic,
Hydrobromic,	Sulphocarbolic,
Hydrochloric,	Sulphuric,
Hydrocyanic,	Sulphurous,
Hydrosulphuric,	Tannic,
Hypophosphorous,	Tartaric,
Lactic,	Valerianic.

Acetic Acid.—Formula (Glacial), $HC_2H_3O_2$. Molecular Weight, 59.86.

Synonym.—Acidum aceticum.

Acetic acid is formed when fruits and saccharine substances pass through the acetic fermentation, vinegar being impure dilute acetic acid. Acetic acid exists in the juices of many trees and plants, and results from the destructive distillation of wood and other organic substances.

Glacial Acetic Acid is the most concentrated form of this substance. It is colorless, liquid at a temperature above 16° C. (60° F.), very mobile, and has a violently penetrating and pungent odor. Its specific gravity is 1.063, boiling point 120° C. (248° F.) (Fowne), and freezing point 15.5° C. (59.9° F.) It is made by distilling finely powdered anhydrous acetate of sodium with three times its weight of concentrated sulphuric acid, according to the following equation:—

$$NaC_2H_3O_2 + H_2SO_4 = NaHSO_4 + HC_2H_3O_2.$$

It mixes with ether, water or alcohol, in all proportions.

Dissolves many resins, oils, camphor and other substances which are insoluble in water and in other menstrua. It should be preserved in glass-stoppered vials.

Acetic Acid (commercial) is of the specific gravity 1.047. It may be prepared from the pyroligneous impregnated for making glacial acetic acid, water being added to the contents of the retort before distillation. It is chemically produced by fermentation, and is present when vinegar impregnates the acetic acid is thus made. The larger portion of the acetic acid, however, that is used in the arts, and even in medicine, is the result of the destructive distillation of wood or pure crystals acid. This latter form is often very impure, and then is positively objectionable as a medicine. The salts of acetic acid contain the radical C_2H_3O, and are acetates.

Properties and Tests.—Acetic acid should be of specific gravity 1.047, and 2 fl. 10 grains diluted with half an ounce of water will exactly neutralize 31 grains of crystallized carbonate of potassium. It should have a sharp, penetrating odor free from empyreumatic contamination, and when 1 part with seven parts of water should yield an agreeable odor of vinegar, with the absence of a smoky smell. The solution should not decolorize when it is briskly tinged with permanganate of potassium, even if previously neutralized with carbonate of potassium. If it produces a precipitate with solution of nitrate of barium, sulphuric acid or a soluble sulphate is present.† If it produces a precipitate with a solution of nitrate of silver, hydrochloric acid or a soluble chloride is present. If a black precipitate results when a current of sulphide of hydrogen is passed through acetic acid, either *copper* or *lead* is present. The other impurities likely to be found in acetic acid are unimportant. Acetic acid is

† It must be borne in mind that vinegar which is made from fermentation contains *soluble sulphates* and will give precipitates with solutions of barium salts. The writer is cognizant of one instance where a large amount of vinegar was returned to a manufacturer when it was perfectly free from sulphuric acid, the analyst having taken it for granted that the production of barium sulphate indicated *free* sulphuric acid.

sometimes designated as *acetic acid number* 8, a name doubtless, derived from the fact that one part of acetic acid mixed with seven parts of distilled water forms the officinal

Diluted Acetic Acid, often called distilled vinegar. This has a specific gravity of 1.006, and one hundred grains will neutralize 7 6-10 grains of crystallized bicarbonate of potassium. *Acetum* of the Pharmacopœia is properly *distilled vinegar,* and is made by the distillation of vinegar.

Arsenic Acid.—Formula, H_3AsO_4. Molecular weight, 141.74.

If powdered arsenious acid of commerce As_2O_3, be dissolved in hot hydrochloric acid and nitric acid be gradually added until red vapors cease to appear, and then the solution be carefully evaporated to dryness and the residuum heated to redness, *arsenic oxide,* As_2O_5, is produced. This dissolves in water to form a solution which deposits, upon evaporation, crystals of *arsenic acid* $2H_3AsO_4.H_2O$, which when heated to 100° C. (212° F.) parts with its water of crystallization, H_3AsO_4 remaining. Arsenic acid is white, poisonous, and is used considerably in the arts as an oxidizing agent. The salts of arsenic acid are called *arsenates.* Arsenic acid yields a precipitate with solution of nitrate of silver, of a reddish brown color. For the detection of this substance see arsenic.

Arsenious Acid.—Formula, H_3AsO_3. Molecular weight, 125.78.

The above is the true arsenious acid, although the pharmacopœia recognizes under the name *arsenious acid* the white arsenic of commerce, which in reality is arsenious anhydrid, As_2O_3. When arsenious anhydrid is dissolved in water an acid solution, said to have the above composition (H_3AsO_3), is formed. It is very unstable and can not be evaporated without decomposition.

Arsenious Acid of Commerce.—Formula, As_2O_3. Molecular weight, 197.68.

Synonyms.—Arsenic. Arsenious oxide. Arsenious anhydrid. White oxide of arsenic. White arsenic.

When the element arsenic is heated in contact with air it unites, forming *arsenious oxide*, As_2O_3, generally known as arsenic, being the *arsenious acid* of our pharmacopœia. It is manufactured in immense quantities as a by-product in roasting the ores of iron, nickel, and cobalt, which contain appreciable amounts of arsenides of these elements. Arsenious acid is in the form of translucent or porcelain-like fragments. When freshly prepared it is translucent and glassy, but by age becomes opaque white from molecular alteration. It is very heavy, odorless, and is said to have a sweetish taste. The crude form dissolves in about nine parts of hot water, the glassy in less proportion. It dissolves freely in hot hydrochloric acid, being deposited upon cooling in the form of octahedral crystals. Alkaline solutions dissolve it, *arsenites* being produced; these salts are unstable. The arsenite of potassium is used in solution as a medicine, under the name Fowler's Solution.

Identification.— See also detection for arsenic. Sulphide of hydrogen throws down from a solution of arsenious acid acidulated with hydrochloric acid, a bright yellow precipitate of sulphide of arsenic. This precipitate is dissolved by ammonia and reprecipitated by acids. Sulphate of copper precipitates solution of arsenious acid in alkaline solution as a green precipitate. This is soluble in ammonia.

Marsh's Test is dependent upon the fact that metallic arsenic unites with hydrogen to form arseniuretted hydrogen, from the flame of which the metallic arsenic is deposited as a brilliant film upon a cold porcelain or glass plate. This result is also obtained by means of antimony, but the two films may be easily distinguished. Arsenious acid and all preparations of this acid are frightful poisons. The usual antidotes for arsenious acid are hydrated peroxide of iron in a fresh, pulpy state; or calcined magnesia. The solution of oxychloride of iron known as *dialyzed iron* has also been recommended as an antidote in connection with bicarbonate of sodium, the soda serving to precipitate the ferric hydrate.

Benzoic Acid.—Formula, $HC_7H_5O_2$. Molecular weight, 121.71.

Benzoic acid exists in many balsams and gum-resins. It is abundantly obtained from benzoin by simply heating, and is the result of the oxidation of certain organic bodies, as toluene. It is formed when hippuric acid is boiled with strong hydrochloric acid, and as the urine of herbiverous animals contains hippuric acid in a large amount, it may be readily prepared from the urine of horses and cows.

Preparation.—Spread a thin layer of benzoin upon the bottom of an iron vessel, and invert over it a paper cap. Upon applying heat the acid will sublime and crystallize in silky needles, condensing upon the inner surface of the paper. This is true medicinal benzoic acid, and should be employed in medicine at the present day, although it is largely superseded by that made from urine, which is cheaper.

Properties.—Benzoic acid from benzoin, is in the form of white silky needles or plates, and is of a pleasant balsamic odor (odorless when pure), due to the presence of traces of essential oil. It evaporates freely, melts at 121° C. (249.8 F.), dissolves in about 200 parts of cold or 25 parts of boiling water, in 10 parts of glycerin, and is very soluble in most essential oils, alcohol or ether. It forms salts with alkalies mostly quite soluble; these are known as *benzoates.*

Examination.—This acid is often made from urine, and when such is the case has generally a powerful, pungent and disagreeable odor, the sense of smell alone being usually sufficient to show the substitution. When made from urine it is generally in larger crystals than from benzoin, and of a satiny luster. According to Hoffmann, hippuric acid is detected by heating a mixture of 20 grains of the acid, 50 grains of hydrate of potassium and 50 drops of water, in a test tube, hippuric acid being indicated if the mixture gives off the odor of ammonia or produces white vapors when a glass rod moistened with acetic acid is held over the mouth of the tube, or if under like conditions it blackens a strip of unsized paper previously moistened

with a solution of mercurous nitrate. *Cinnamic acid* may be present in benzoic acid produced from gum benzoin, but is of little consequence from a therapeutical point. It may be detected by heating a concentrated aqueous solution of benzoic acid with permanganate of potassium, when, if cinnamic acid is present, an odor of oil of bitter almonds will be evolved.

Boric Acid.—Formula, H_3BO_3. Molecular weight, 61.88.

This is also known under the name *boracic acid* and occurs free in the waters of certain volcanic districts of Tuscany, from which it is obtained by evaporation and crystallization. It may be easily produced artificially by dissolving borate of sodium (borax) to saturation in boiling water and adding an excess of hydrochloric acid, when upon cooling an abundance of boric acid will separate in crystalline form. Boric acid is translucent, of a beautiful satin-like, pearly appearance, resembling fish scales. It is soluble in three parts of boiling water, and dissolves considerably in alcohol, the solution in alcohol burning with a green flame. It is almost tasteless and slightly reddens blue litmus paper. Boric acid is scarcely liable to adulteration, but has been used for the purpose of adulterating santonin. One specimen of santonin (?) examined by the writer contained 85 per cent. of boric acid. It can be readily detected by means of chloroform which freely dissolves the santonin without affecting boric acid.

Carbazotic Acid.—Formula, $HC_6H_2(NO_2)_3O$. Molecular weight, 228.57.

Synonyms.—Picric acid, Tri-nitro-phenic acid.

This is made by acting upon carbolic acid with nitric acid. It may also be prepared in the same manner from indigo, silk, wool and many other organic bodies, especially resins and gums. The most economical process is from carbolic acid, but the reaction is often violent and even dangerous; therefore, as carbazotic acid is made in large amounts by manufacturers, with facilities for the purpose, it is better to purchase it for medicinal use and purify if

necessary. Picric acid is in the form of bright lemon yellow crystals, melts at 122.5° C. (252.5° F.), dissolves slightly in cold water, to a considerable extent in boiling water, and quite freely in alcohol. It is very bitter to the taste, stains organic substances yellow, and for this reason caution must be observed in handling it, as the color produced can not be removed from the nails and hair, a new growth being necessary. Carbazotic acid unites with alkalies to form salts, the ammonium salt being quite soluble and considerably employed in medicine. The salts of carbazotic acid are called *carbazotates* (or picrates) all the soluble salts being exceedingly bitter. Carbazotate of potassium is a dangerous compound, exploding with violence when heated. Carbazotic acid unites with alkaloids to form salts which are mostly insoluble. Carbazotic acid is largely employed in the arts as a dye-stuff. It imparts a fine yellow color.

Carbolic Acid.—Formula, HC_6H_5O. Molecular weight, 93.78.

Synonyms.—Phenol. Phenic acid. Phenyl alcohol. Phenylic hydrate. Coal-tar creasote.

Carbolic acid is practically obtained from coal-tar, although it may be made by way of experiment by heating salicylic acid quickly in a glass retort, carbolic acid distilling. When crude coal-tar is agitated with milk of lime and permitted to stand some days, the aqueous liquid which rises, if decanted and neutralized with hydrochloric acid, will yield an oily substance. This is *crude* carbolic acid, and may be partially purified by subsequent fractional distillations, although the preparation of *pure* carbolic acid is most tedious and difficult, seldom being accomplished excepting on a manufacturing scale.

Carbolic acid is usually met with in crystalline masses, colorless, of a strong penetrating odor and caustic burning taste. The crystals melt at about 42° C. (108° F.), producing a colorless liquid that may be often reduced much below this temperature without crystallizing. Under such circumstances a crystal of carbolic acid added will cause the mass to solidify, and even a sharp blow upon the

container will accomplish the same result. The crystals absorb moisture upon exposure, and form a hydrate ($HC_6H_5O.H_2O$), which melts at a much lower temperature. Carbolic acid of commerce is often of a reddish color, and frequently becomes red by age. If specimens of such are melted and poured into a glass funnel stopped at the exit, and crystallized by reducing the temperature, then placed in a moderately warm room and the cork removed, as the acid slowly melts the red color passes away with the first liquid, leaving the mass of crystals pure white. Carbolic acid is an antiseptic and is much employed as a disinfectant. It is soluble in glycerin, and carbolic acid mixed with glycerin, bulk for bulk, will mix with water in all proportions. When permitted to come into contact with the skin it acts as a caustic, turning the cuticle white, benumbing the part exposed to its action, and producing a painful sore. Olive or linseed oil applied at once will generally neutralize its serious effects. Carbolic acid in overdoses is an active poison, burning the throat and stomach, death often following in a short time. The antidotes are oils, milk, or white of egg, to be freely administered. Evacuate the stomach as soon as possible.

Carbonic Acid.—Formula, H_2CO_3. Molecular weight, 61.85.

When carbon is burned in air or in oxygen a colorless gas results, composed of two atoms of oxygen united with one atom of carbon. This gas is *carbon dioxide*, CO_2, often called carbonic acid gas, and sometimes improperly designated as carbonic acid. When this gas passes into cold water one molecule unites with a molecule of water to form a molecule of true *carbonic acid*, thus:—

$$CO_2 + H_2O = H_2CO_3.$$

Carbonic acid is usually made, however, by decomposing limestone, $CaCO_3$, by means of sulphuric acid.

$$CaCO_3 + H_2SO_4 = CaSO_4 + H_2CO_3.$$

In this case, however, the carbonic acid decomposes into carbon dioxide and water as follows:—

$$H_2CO_3 = H_2O + CO_2.$$

It is therefore necessary to pass the carbon dioxide, as it is formed, through a pipe and into cold water, when combination again ensues in proportion to the pressure. This is practically the process employed in making effervescing waters. Carbonic acid is always dissolved in water, not being known free, and is found in the waters of many springs. It imparts a distinct acid reaction, changing blue litmus to red, and has a sharp agreeable taste. If carbon dioxide be passed into lime water a turbidity results, consequent upon the formation of carbonate of calcium. If diluted acids are poured upon a carbonate, decomposition ensues, carbon dioxide being liberated. The salts of carbonic acid are called *carbonates.*

Chloric Acid.—Formula, $HClO_3$. Molecular weight, 84.25.

This acid has never been obtained in a pure state owing to its ready decomposition. It has been produced as a syrupy liquid by decomposing chlorate of potassium with a solution of hydrofluosilicic acid, and separating the excess of hydrofluosilicic acid with silica, then evaporating. It is very explosive if permitted to come into contact with many organic bodies, and will even inflame paper, so easily is it deoxidized. This acid is never used in medicine, but the salts of it are known as chlorates and the chlorate of potassium ($KClO_3$) is consumed extensively.

Chromic Acid.—Formula, H_2CrO_4. Molecular weight, 118.24.

When one and one-half parts of concentrated sulphuric acid are slowly, and with constant stirring, poured into one part of a cold saturated solution of bichromate of potassium, and the mixture permitted to cool, beautiful red crystals of trioxide of chromium (CrO_3) separate. These crystals drained upon a soft brick in dry atmosphere, and quickly placed in glass stoppered bottles, constitute the so-called chromic acid of the U. S. P., although to form the true chromic acid it is necessary to dissolve the above preparation in water, thus:—

$$CrO_3 + H_2O = H_2CrO_4.$$

Trioxide of chromium (officinal chromic acid) is of a red color and deliquesces in a moist atmosphere, forming a solution of the chromic acid. It is a powerful oxidizing agent and should never be mixed with organic bodies. It strongly decomposes flesh and by some is considered among the most valuable of caustics, being extensively employed to remove warts and similar growths. The salts of chromic acid are called *chromates*.

Chrysophanic Acid.—Formula, $C_{15}H_{10}O_4$. Molecular weight, 254.12.

Synonyms.—Parietin. Rheic. Rhabarbaric acid. Rumicin, etc.

This substance was discovered in the lichen *Parmelia parietina*, and has been prepared from rhubarb, yellow dock root and senna pods. It occurs in the form of yellow needles or powder, is tasteless and odorless. It slightly dissolves in boiling water, alcohol and ether, freely in warm benzol, forming a yellow solution. Alkalies dissolve it freely, yielding deep red colored solutions. Some doubt has recently been expressed regarding the presence of chrysophanic acid in senna, and it is now accepted that aloin contains a substance which by oxidation produces chrysophanic acid.

Citric Acid.—Formula, $HC_6H_7O_7$. Molecular weight, 191.54. Crystals, $2[H_3C_6H_5O_7 \cdot H_2O]$. Molecular weight, 419.50.

Citric acid is found in the juice of many fruits, although the acid of commerce is made from the juice of lemons and limes. If lemon juice be boiled and then saturated with chalk or lime, citrate of calcium $Ca_3(C_6H_5O_7)_2$ will precipitate and must be strained from the hot liquid, for it will dissolve upon cooling. The addition of water and sulphuric acid now forms sulphate of calcium and citric acid according to the following equation:—

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[1] Lemons and limes are from South America, see supplement to King's American Dispensatory by King and Lloyd, and the National Dispensatory by Stillé and Maisch.

$$\underset{\text{Citrate of Calcium.}}{Ca_3 2C_6H_5O_7} + \underset{\text{Sulphuric acid.}}{3H_2SO_4} = \underset{\text{Citric acid.}}{2H_3C_6H_5O_7} + \underset{\text{Sulphate of Calcium.}}{3CaSO_4}.$$

By evaporating the filtered liquid, crystals of citric acid may be obtained after exposure to low temperature. Citric acid is consumed in medicine in large quantities. It forms transparent crystals, very soluble in hot water, freely in cold water, slightly in alcohol, and insoluble in ether. A peculiarity of citric acid is the fact that its calcium salt is quite soluble in cold water, and is scarcely dissolved by hot water. It is very sour to the taste, is not poisonous, and forms salts which are known as *citrates*, several of its salts being of value as therapeutical agents.

Formic Acid.—Formula, $HCHO_2$. Molecular weight, 45.89.

Formic acid derives its name from the red ant (Formica rufa), from which it was first obtained. These ants, and some other insects, contain considerable amounts of formic acid, as do some plants, like the nettle. Formic acid is made artificially by heating equal weights of oxalic acid and glycerin together for about fifteen hours, at a temperature slightly above that at which water boils, and afterwards distilling the residue with water. The glycerin is not altered but the oxalic acid decomposes, forming formic acid and carbon dioxide, thus:—

$$\underset{\text{Oxalic acid.}}{H_2C_2O_4} = \underset{\text{Formic acid.}}{HCHO_2} + \underset{\text{Carbon dioxide.}}{CO_2}.$$

Formic acid is a clear liquid of specific gravity 1.235, crystallizing, if pure, below the freezing point of water. It attacks the skin and has a sour taste and odor. It reduces the salts of the noble metals (gold, etc.) and of mercury. The salts of formic acid are known as *formates*.

Gallic Acid.— Formula, $H_3C_7H_3O_5,H_2O$. Molecular weight, 187.55.

Gallic acid is found in some vegetable substances, but not in sufficient amount to repay extraction. It is produced from oak-galls, by the decomposition of the tannic acid they contain so abundantly. This is accomplished

by moistening powdered gall-nuts with water and exposing to a warm temperature some weeks, occasionally adding water to replace that which evaporates. This pulp is then exhausted with boiling water, filtered and cooled, and the crystals of impure gallic acid are again dissolved in water, and if necessary, are decolorized by means of animal charcoal, when upon cooling pure gallic acid crystallizes. There has been much dispute regarding the chemistry of this conversion of tannic into gallic acid, but, perhaps, the most acceptable explanation of the reaction is afforded by Messrs. H. Schiff, Sac and J. Löwe, upon the theory that tannic acid is an *anhydrid* of gallic acid, and that the conversion of tannic acid into gallic acid may be explained as follows:—

$$\underset{\text{Tannic acid.}}{C_{14}H_{10}O_9} + \underset{\text{Water.}}{H_2O} = \underset{\text{Two molecules of gallic acid.}}{2C_7H_6O_5}.$$

Gallic acid crystallizes in small needles of a light fawn color, often approaching white. It dissolves in three parts of boiling water and in 100 parts of cold water; dissolves freely in hot alcohol; sparingly in ether and in glycerin. It does not precipitate solution of gelatin, (distinction from tannic acid), but with solution of acetate of copper produces a precipitate insoluble in water, but very soluble in ammonia water.

Hydriodic Acid.—Formula, HI. Molecular weight, 127.53.

This acid (gas) may be made by allowing iodine and phosphorus to react upon each other under water.

Preparation.—Mix one part of amorphous phosphorus with fifteen parts of water in a tubulated retort connected by means of a rubber tube with a glass tube, the exit of which is plunged into water. Now surround the retort with ice water and gradually add iodine through the tubulus of the retort until twenty parts of iodine have been used. When the gas ceases to pass, warm the retort gently until the reaction ceases. The gas which is evolved is *hydriodic acid* and the solution of it in the water is the hydriodic acid of commerce. This gas may be

made otherwise, as by heating iodide of potassium with phosphoric acid in solution, or by passing sulphide of hydrogen into a mixture of powdered iodine and water.

Hydriodic acid gas is colorless, strongly acid in reaction, fumes in the atmosphere, and is very suffocating if inhaled. It liquefies at 0° C, (32° F.) under a pressure of four atmospheres (Faraday,) and freezes at —55° C. (—67° F.) to a mass resembling ice. The specific gravity is 4.37, air being the standard of unity.

The solution of hydriodic acid resembles hydrochloric and hydrobromic acids. It decomposes by exposure to the atmosphere, becoming reddish-brown, from liberation of iodine which remains dissolved in the acid solution. Hydriodic acid forms salts known as *iodides*.

Hydrobromic Acid.—Formula, HBr. Molecular weight 80.75.

Hydrobromic acid (bromide of hydrogen) contains eighty parts of bromine, by weight, to eighty-one parts of the acid. This is a larger percentage of bromine than is afforded by any other salt of bromine. It may be made by several indirect processes and by the direct combination of its elements. A few drops of bromine should be placed in the bottom of a long *cold* test-tube and the tube, filled to the height of two inches with small fragments of glass upon which is placed a few grains of phosphorus, and above this two inches of broken glass as before, the upper portion of the last layer of glass being wetted. Now apply a gentle heat, the bromine vaporizes, unites with the phosphorus, and forms bromide of phosphorus (PB_5); this is decomposed by the water upon the wet glass, forming phosphoric and phosphorous acids, and hydrobromic acid gas, which latter may be led through a glass tube and condensed. The hydrobromic acid used in medicine is a solution of this gas in water, and is made by passing the gas into cold water. The foregoing experiment can not be conducted upon a large scale, as the reaction between the bromine vapor and phosphorus is violent and dangerous (not more than ten grains of bromine and phosphorus should be experimented with), there-

fore other processes are used in practice, usually by distilling a mixture of bromide of potassium solution with sulphuric acid, and passing the gas into water, thus:—

$$\underset{\text{Bromide of potassium.}}{2KBr} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Hydrobromic acid.}}{2HBr} + \underset{\text{Sulphate of potassium.}}{K_2SO_4}.$$

Hydrobromic acid gas is colorless, pungent and irritating, and fumes in the atmosphere. Hydrobromic acid of commerce should contain 34 per cent. of anhydrous acid, and 95 parts should saturate 20 parts of dry carbonate of potassium.† If sulphuric acid be present it will give a precipitate with solution of chloride of barium. Hydrobromic acid should perfectly evaporate. The salts of hydrobromic acid are *bromides*, some being extensively used in medicine.

Hydrochloric Acid.—Formula, HCl. Molecular weight, 36 37.

Synonyms.—Muriatic acid gas. Hydrochloric acid gas.

Hydrochloric acid gas is composed of equal volumes of hydrogen and chlorine united by chemism, thus:—

$$H_2 + Cl_2 = 2HCl.$$

It is made for use in the arts by heating in a retort, equal parts of sulphuric acid and chloride of sodium. The reaction is expressed as follows:—

$$\underset{\text{Sodium chloride.}}{2NaCl} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Sodium sulphate.}}{Na_2SO_4} + \underset{\text{Hydrochloric acid.}}{2HCl}.$$

In this case normal sulphate of sodium also is formed. At a temperature of less than 200° C. (392° F.) only half the hydrochloric acid is liberated and the result is a residuum of acid sulphate of sodium, thus:—

$$\underset{\text{Chloride of sodium.}}{2NaCl} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Acid sulphate of sodium.}}{NaHSO_4} + \underset{\text{Chloride of sodium.}}{NaCl} + \underset{\text{Hydrochloric acid.}}{HCl}.$$

When the temperature rises above a certain point the molecule of free sodium chloride and the molecule of acid

†Report of the committee appointed by the American Pharmaceutical Association, for the revision of the U. S. P.

sulphate of sodium are mutually decomposed, resulting in the formation of another molecule of hydrochloric acid:—

$$\underset{\text{Acid Sulphate of sodium.}}{NaHSO_4} + \underset{\text{Chloride of sodium.}}{NaCl} = \underset{\text{Sulphate of sodium.}}{Na_2SO_4} + \underset{\text{Hydrochloric acid.}}{HCl}.$$

Hydrochloric acid gas is pungent, irrespirable and extinguishes flame. A pressure of 40 atmospheres at 10° C. (50°F.), condenses it into a transparent liquid. It reddens blue litmus paper, and when permitted to escape into the atmosphere, seizes upon the moisture present with such avidity as to form heavy fumes. When conducted into water it is dissolved, a solution of 30 to 33 per cent. of the gas in water constituting

Officinal Muriatic (Hydrochloric) Acid.

Synonyms.—Acidum Hydrochloricum. [Spirit of salt.)

This has the specific gravity of 1.160 and should be perfectly transparent, although often yellowish from the presence of impurities. It has a suffocating odor and a sharp, sour, corrosive taste. It fumes in the air and evolves chlorine when heated with dioxide of manganese.

Tests.—If a residue remains upon evaporation, fixed impurities are present, likely salts of sodium. If diluted with four times its bulk of water sulphide of hydrogen should not produce a turbidity (absence of metals); with chloride of barium a precipitate will follow if sulphuric acid or a soluble sulphate is present. Arsenic is sometimes present in hydrochloric acid and may be detected according to the German Pharmacopœia as follows: "Put a few small pieces of pure zinc into a rather long test-tube, and introduce the hydrochloric acid, diluted with two parts of water, which should fill about one-tenth part of the tube; into the upper portion of the tube is placed a pellet of cotton-wool saturated with a solution of acetate of lead, and the mouth of the vessel is covered with a piece of white bibulous paper moistened with a solution of nitrate of silver. After an active evolution of gas, for half an hour, in the absence of sulphurous or arsenious acid, neither the cotton (sulphurous acid) nor the paper (arsenious acid) should be blackened." Hydrochlo-

ric acid in concentrated form is a corrosive poison, only less violent in its action than sulphuric and nitric acids. The effects upon the tissues are similar, and the antidotes are the same as those named for sulphuric and nitric acids.

Diluted Muriatic Acid of the U. S. P. is made by mixing four troy ounces of hydrochloric acid with sufficient distilled water to make sixteen fluid ounces; this has a specific gravity of 1.038. The salts of hydrochloric acid are called properly *chlorides*, although they are familiarly known in commerce as muriates or even hydrochlorates.

Hydrocyanic Acid.—Formula, HCN (or HCy). Molecular weight, 26.98.

Synonym.—Prussic acid.

When the salt known as cyanide of mercury is heated it decomposes, the metallic mercury being liberated and a gas escaping which has the composition C_2N_2 (cyanogen gas). This gas contains, in each molecule, two molecules of the acidulous radical (CN) of hydrocyanic acid. This acidulous radical is often called cyanogen and represented by Cy, therefore, we have the formula HCy, for hydrocyanic acid meaning cyanide of hydrogen. Anhydrous hydrocyanic acid may be obtained by passing dry sulphide of hydrogen over cyanide of mercury heated gently in a glass tube, the gas being led into a freezing mixture where it condenses as a colorless liquid. This is very volatile, possesses a feeble acid reaction and a powerful odor resembling bitter almond oil or peach kernels. It is one of the most fearful of known poisons, the inhalation of its vapor being almost instantaneously fatal to animal life; therefore, the utmost care must be used in experimenting with it,† few being willing to run the risk of its production. When largely diluted it constitutes the—

Officinal Diluted Hydrocyanic Acid —This contains two (2) per cent. of anhydrous acid,‡ and is made by

† Scheele who discovered this acid died suddenly while conducting some investigation, and it is supposed that he was poisoned by the vapor of the acid.

‡ Scheele's acid contains from 3 to 5 per cent.

distilling a mixture of ferrocyanide of potassium, sulphuric acid and water, the reaction being mostly as follows:—

$$2K_4FeCy_6 + 3H_2SO_4 = 6HCy + K_2Fe_2Cy_6 + 3K_2SO_4.$$

Ferro cyanide of potassium.	Sulphuric acid.	Hydrocyanic acid.	Potassio-ferrous ferro-cyanide.	Potassium sulphate.

The distillation must be conducted slowly owing to the tendency of the contents of the retort to "bump," often with sufficient violence to break the retort. This acid is liable to decompose and turn black even though securely sealed and protected from the light. This trouble, according to experiments, may be overcome by adding alcohol instead of water to the distillate, for the distillate is concentrated and must be reduced to the percentage required. Hydrocyanic acid may be obtained by the decomposition of several cyanides. The U. S. P. recommends that 50½ grains of cyanide of silver be added to a mixture of 41 grains of hydrochloric acid and one fluid ounce of water, and after being well shaken together, permitted to settle, when the clear solution may be decanted from the chloride of silver which precipitates.

Tests.—In cases of poisoning by hydrocyanic acid the characteristic odor of the acid is said to be usually distinct and unmistakable about the mouth and nostrils of the deceased, often perceptibly pervading the entire room. This acid is detected chemically by several simple and unmistakable reactions. With solution of nitrate of silver it throws down a dense white precipitate, which blackens by exposure to light, but not so rapidly as precipitated chloride of silver, the odor of the acid disappearing if the silver solution be in excess; it dissolves in boiling nitric acid, (chloride of silver does not), and decomposes at a red heat, yielding *metallic silver;* chloride under the same condition simply fusing. Scheele's test is to mix the suspected liquid with a few drops of solution of ferrous sulphate and an excess of hydrate of potassium, then stir well in a shallow vessel for ten or fifteen minutes; this will form *Prussian blue* mixed with precipitated oxide of iron; the addition now of hydrochloric acid will dissolve the oxide of iron, and if hydrocyanic acid is present will leave a blue insoluble precipitate (Prussian blue). It

must be remembered that hydrocyanic acid rapidly decomposes in the stomach, and it is said is completely destroyed within twenty-four hours after death. Therefore the examination of the contents of the stomach can not be deferred as with most other poisons.

Diluted Hydrocyanic Acid is often used as a poison. Although necessarily weaker than the anhydrous it is violent and rapid in its action, the vapor being sufficient to occasion serious effects The writer inadvertently inhaled the vapor while distilling the acid, and was suddenly overcome by a depression, stupor, and inability to stand. There was no pain or sensation of warning; fresh air immediately removed the dizziness. The after effects were not unpleasant.

Antidote.—When prussic acid has been taken in sufficient amount, its action is so rapid as to preclude the use of antidotes of a chemical nature, authenticated instances showing that it poisons even in the act of swallowing. Immediately pour a stream of cold water from the height of five or six feet upon the spine and back of the head of the patient; and artificial respiration may be resorted to if there is any hope, and weak ammonia applied to the nostrils. If there has been any considerable amount swallowed there is no hope. Death will be almost instantaneous. Physicians and druggists can not be too cautious regarding this substance so antagonistic to animal life. It resembles distilled water in appearance, has an agreeable peach-pit odor, is apparently harmless, and is liable to be taken by children or others without exciting suspicion. The salts of hydrocyanic acid are called *cyanides* and the cyanide of potassium is extremely poisonous (see page 178), often being used with suicidal intention, and the attempt is usually a success.

Hydrosulphuric Acid.—Formula, H_2S. Molecular weight, 33 98.

Synonyms.—Sulphide of hydrogen. Sulphuretted hydrogen.

This is a gas at ordinary temperatures, colorless, heavier than air, three volumes dissolving in one volume of water at ordinary temperature. It condenses at —74° C.(—101.2°

F.) under the pressure of seventeen atmospheres into a colorless liquid which freezes at —85° C. (—121° F.) to a mass resembling ice. It has a disgusting odor, resembling rotten eggs, burns with a pale blue flame, and reacts slightly as an acid, reddening blue litmus. It is found in the waters of so-called sulphur springs, and imparts to them the odor of putrid eggs.

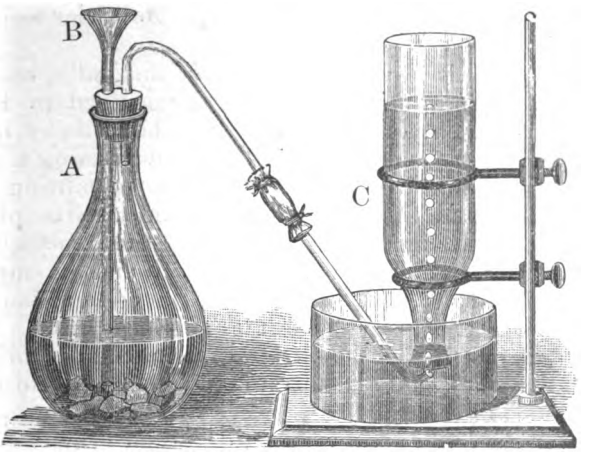

FIG. 48.

Sulphide of hydrogen may be readily made as follows, (fig. 48): Into the chemical flask A, place a few pieces of sulphide of iron, and close the mouth of the flask with a stopper fitted with the funnel tube B, and a bent glass tube, which latter leads to the pneumatic trough C. Now pour through the funnel tube a mixture of sulphuric acid one part, and three parts of water; the sulphide of hydrogen (gas) will be at once and regularly evolved, until the sulphide of iron is decomposed, or the sulphuric acid is saturated; thus:—

$$\underset{\text{Sulphide of iron.}}{FeS} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Sulphide of hydrogen.}}{H_2S} + \underset{\text{Sulphate of iron.}}{FeSO_4}.$$

Sulphide of hydrogen is not used as a medicine if we may except natural mineral waters. It is exceedingly

deleterious and even poisonous if inhaled in large amounts, and therefore should be prepared in the open atmosphere or under a flue with a good draft. Faraday states that an atmosphere which contains $\frac{1}{1500}$ of this gas will destroy the life of birds, and that dogs die when it contains only $\frac{1}{800}$ part. Sulphide of hydrogen forms *sulphides*.

Lactic Acid.—Formula, $H_2C_3H_4O_3$. Molecular weight, 89.79.

Lactic acid is the principal acid of sour milk, and is formed by fermentation of the sugar contained in milk through the agency of caseine. It may be made by mixing sugar, cheese and chalk together, and forming a thin paste of the mixture with water; then permitting the mixture to stand three or four weeks in a warm place, occasionally stirring and supplying water to replace that which evaporates. The mass is strained and the impure crystals of *lactate of calcium* recrystallized from boiling water, and decomposed by dilute sulphuric acid. The mixture is filtered and the filtrate evaporated and purified by solution in alcohol and filtration. The lactic acid may be obtained of a syrupy consistence by evaporation.

Lactic Acid is a syrupy liquid of specific gravity 1.212 (U. S. P.) almost odorless, colorless, and is very sour to the taste. It mixes with water, glycerin, or alcohol, in all proportions. The most likely impurities are calcium salts and sulphuric acid. The former can be shown by the turbidness which results from adding oxalate of ammonium to the acid diluted with four times its bulk of water. Sulphuric acid, if present, will be indicated by the turbidness which results by the addition of solution of chloride of barium to diluted lactic acid. If the acid be of an offensive odor, or of a dark color, organic impurities are present. The salts of lactic acid are known as *lactates*, and a discrimination must be made between these salts and mixtures of milk sugar with other substances which are sometimes improperly called "lactates."

Malic Acid.—Formula, $H_3C_4H_3O_5$. Molecular weight, 133.68.

Malic acid is found in the juice of many fruits, such as green apples, gooseberries, currents, etc., and very abundantly in the stalks of garden rhubarb, existing as acid malate of calcium, $CaH_2.2C_4H_3O_5$. It may be prepared by digesting the juice of rhubarb stalks with lime, filtering the mixture and evaporating the filtrate to the consistence of a thin syrup. This is mixed with eight times its bulk of alcohol, the precipitate collected in a muslin strainer and pressed. The malate of calcium thus obtained is decomposed with dilute sulphuric acid, filtered again, and afterward the solution is evaporated to crystallization, towards the last at a low temperature beneath the receiver of an air pump.

Malic acid is deliquescent, and from this reason perhaps is not used in medicine in the pure form. The German Pharmacopœia recognizes both a ferrated tincture of sour apples, and a ferrated extract of sour apples. In each case the preparation contains *malate* of *iron* in proportion to the malic acid present in the fresh apples. *Acetate* of *iron* is also present and results from fermentation of the juice. Malic acid forms salts which are known as *malates*.

+ **Nitric Acid** —Formula, HNO_3. Molecular weight, 62.89

Synonyms.—Aqua-fortis.

Nitric acid may be produced by passing electric sparks through moist atmospheric air; also by burning many substances in a mixture of oxygen and nitrogen; by distilling a mixture of saltpetre, sulphate of copper and alum;† by passing an electric spark through a mixture of oxygen and hydrogen and atmospheric air, and by the decomposition of various nitrates. It is made practically by distilling a mixture of equal weights of sulphuric acid and nitrate of potassium, or nitrate of sodium, the distillate being collected in a cooled receiver, nitric acid

† Gober prepared it in this manner.

and acid sulphate of potassium (bisulphate of potassium) being formed, thus:—

$$\underset{\text{Nitrate of potassium.}}{KNO_3} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Nitric acid.}}{HNO_3} + \underset{\text{Acid sulphate of potassium.}}{KHSO_4}.$$

As thus obtained we find nitric acid contaminated with *nitrogen peroxide* (NO_2), which gives a yellow color to the commercial product. The specific gravity of nitric acid, U. S. P., is 1.42, it contains 69.96 per cent. of anhydrous nitric acid, and it is generally stated that an acid of less or of greater strength than the above changes when boiled (water or nitric acid separating) until it boils regularly, an acid of specific gravity 1.42 then distilling. Roscoe states that the constant boiling point of the pure acid is 120.5° C. (248.9° F.), and that the acid has the specific gravity of 1.414. *Pure nitric acid* is colorless and fumes in the air. It possesses a disagreeable, peculiar odor, and absorbs moisture from the atmosphere. It is of specific gravity 1.559 (Kolbe), is very corrosive and produces serious sores when permitted to come into contact with the skin. The officinal nitric acid (specific gravity 1.42), quickly colors the skin and nails yellow and if sufficiently in contact with the skin, produces penetrating wounds which are slow to heal.

Diluted Nitric Acid, U. S. P., is made by mixing three troy ounces of officinal nitric acid with sufficient distilled water to make sixteen fluid ounces. It has the specific gravity of 1.068.

Tests and Impurities.—If a crystal of ferrous sulphate (green vitriol) be placed in a test tube with a liquid which is suspected of being contaminated with nitric acid, and sulphuric acid be poured directly upon the crystal by means of a pipette a reddish brown or even black color will be produced where the two liquids are in contact, providing that nitric acid be present.

If nitric acid be diluted with five parts (by measure) of water, and a solution of nitrate of silver be added, a white precipitate will fall if *hydrochloric acid* be present. If the diluted acid be mixed with solution of nitrate of barium, a white precipitate will follow if *sulphuric acid*

be present. If a dry residuum remain upon evaporation of nitric acid, fixed impurities are present.

Nitric acid is a corrosive poison, decomposing the tissues when swallowed, and destroying life with the most terrible of sufferings.

Antidotes.—Either calcined or carbonate of magnesium, or solution of one of the carbonated alkalies, embracing carbonate of sodium and bicarbonate of sodium, or lime water, are antidotes, one of which must be given immediately, and followed by either milk, gruel, oils, or mucilaginous drinks. In case the chemical antidotes can not be obtained at once, administer the oils and viscid drinks, which act by enveloping the poison. It must be remembered that nitric acid *destroys* the tissues, and if the acid be concentrated and the dose be large, the membranes of the throat and stomach are actually decomposed at once, and all hope of ultimate recovery must be abandoned.

Nitro-Hydrochloric Acid. Nitro-Muriatic Acid.

This is a mixture of nitric and muriatic acids, made according to the U. S. P. by mixing three parts, by weight, of nitric acid with five parts of hydrochloric acid. This mixture liberates chlorine, and will dissolve the *noble* metals, gold and platinum, hence the old name, *aqua regia*, from this property of dissolving gold, the king of metals. There are other products of decomposition of little general interest, inasmuch as the chlorine in nascent form is the active agent. The reaction which results upon mixing the acids, may be explained as follows:—

$$\underset{\text{Nitric acid.}}{2HNO_3} + \underset{\text{Hydrochloric acid.}}{6HCl} = \underset{\text{Chloronitric acid gas.}}{2NOCl_2} + \underset{\text{Water.}}{4H_2O} + \underset{\text{Chlorine.}}{Cl_2}.$$

Nitro-hydrochloric acid should be mixed twenty four hours before use. It is a corrosive poison, and the treatment and antidotes are the same as those recommended for poisoning by nitric acid.

Diluted Nitro-Muriatic Acid, U. S. P., is made by mixing one and one-half ounces of nitric acid with two and one-half ounces muriatic acid, permitting the mixture to

remain twenty-four hours, then bringing it to the measure of sixteen fluid ounces by means of distilled water. This should be kept in a cool place.

Nitrous Acid.—Formula, HNO_2. Molecular weight, 46.93.

Nitrous acid may be made by passing *nitrogen trioxide* into ice cold water, thus:

$$\underset{\text{Nitrogen trioxide.}}{N_2O_3} + \underset{\text{Water}}{H_2O} = \underset{\text{Nitrous acid.}}{2HNO_2}.$$

This is a beautiful blue liquid, but not permanent, for upon warming, decomposition ensues as follows:

$$\underset{\text{Nitrous acid.}}{3HNO_2} = \underset{\text{Nitric acid.}}{HNO_3} + \underset{\text{Nitric oxide gas}}{2NO} + \underset{\text{Water.}}{H_2O}.$$

When nitrogen trioxide is passed into warm water, nitric acid and nitrogen dioxide result, thus:

$$\underset{\text{Nitrogen trioxide.}}{3N_2O_3} + \underset{\text{Water.}}{H_2O} = \underset{\text{Nitric acid.}}{2HNO_3} + \underset{\text{Nitrogen dioxide.}}{2N_2O_2}.$$

Nitrous acid is not used in medicine in the free state, but *nitrite of ethyl* ($C_2H_5NO_2$), is the chief ethereal ingredient of spirit of nitrous ether. Although nitrous acid is so unstable, the salts of nitrous acid known as *nitrites* are permanent.

Oleic Acid.—Formula, $C_{18}H_{34}O_2$. Molecular weight, 281.38.

Oleic acid is obtained from many liquid oils by saponification and decomposition of the resultant soap by means of tartaric or hydrochloric acid; the separated fatty acids are then digested with oxide of lead some hours at a temperature of 150° F. This produces lead salts of the mixed fatty acids, from which the oleate of lead is dissolved with ether, the solution filtered and then decomposed with hydrochloric acid. The chloride of lead is separated by filtration of the ethereal solution and the *oleic acid* is obtained from this by evaporation of the ether. Oleic acid is tasteless, odorless, and colorless when pure. It is oily, has the specific gravity 0.898 at 19° C.

(66.2° F.), dissolves freely in alcohol and in ether, but is insoluble in water. Oleic acid forms salts, thus— $KC_{18}H_{33}O_2$ is oleate of potassium, and $Pb2C_{18}H_{33}O_2$ is oleate of lead. Soft soap is chiefly oleate of potassium, and hard soap is chiefly oleate of sodium. There are many other compounds used in medicine which contain salts of oleic acid, as oleate of mercury and the various lead plasters and lead ointments.

Oxalic Acid.—Formula, $H_2C_2O_4 2H_2O$. Molecular weight, 125.70.

Oxalic acid is found in the juices of certain plants in the form of oxalates, most especially the oxalis or sorrel grasses. Oxalic acid is made artificially in immense quantities by the action of nitric acid upon sugar, or of alkalies upon sawdust. Oxalic acid crystallizes with two molecules of water, the water being expelled by a gentle heat, anhydrous oxalic acid ($H_2C_2O_4$) remaining. If the heat be increased until decomposition takes place, formic acid and carbon monoxide, carbon dioxide and water result. Crystals of oxalic acid are colorless, transparent, and dissolve in eight parts of water at 15.5° C. (59.9° F.) or in their own weight of hot water. Oxalic acid is very sour and strongly acid in reaction. When solution of oxalic acid or of the soluble oxalates is added to solution of chloride of calcium oxalate of calcium precipitates. This is white, and if the precipitate does not dissolve upon the addition of acetic acid, it is an excellent test for oxalic acid, free or combined, (acetic acid dissolves phosphate of calcium). The so-called salt of sorrel (KHC_2O_4, $H_2C_2O_4,2H_2O$), as sold upon the market for removing ink stains, is often simply oxalic acid. Oxalic acid is a powerful poison.

Treatment and Antidotes.—Make a creamy mixture of prepared chalk and administer it in doses of from a table spoonful to half a teacupful every ten minutes. Solution of ferrous sulphate (green vitriol) may be resorted to if the chalk can not be obtained, (oxalate of iron is insoluble). In all cases evacuate the stomach as quickly as possible by means of a stomach pump. "In after treat-

ment (in the stage of collapse) warmth should be applied and stimulants administered"—*Taylor*. Do not give soluble alkalies or their carbonates, for oxalates of the alkalies are soluble and are equal in poisonous effects to oxalic acid. Oxalic acid acts usually with extreme rapidity, and if fairly absorbed into the system we may consider that death is almost inevitable.

Phosphoric Acid—(Normal or Ortho-phosphoric Acid.) Formula, H_3PO_4. Molecular weight, 97.80.

This may be made by dissolving common phosphorus in diluted nitric acid, or by the action of concentrated nitric acid upon red (amorphous) phosphorus. Also from bone ash, which is chiefly a tri-calcium phosphate. It is not advisable to prepare this acid from ordinary phosphorus unless the operator has proper appliances and the use of steam, and even then care must be exercised. Phosphoric acid is sometimes made by dissolving glacial phosphoric acid in water and then adding nitric acid to the solution and heating until red vapors cease to be evolved (See U. S. P. 1870, p. 71.) This acid should not be used, as glacial phosphoric acid of commerce is very impure. The reaction which takes place between phosphorus and dilute nitric acid, whereby phosphoric acid is produced, may be explained as follows:—

$$\underset{\text{Phosphorus.}}{3P_4} + \underset{\text{Nitric acid.}}{20HNO_3} + \underset{\text{Water.}}{8H_2O} = \underset{\text{Phosphoric acid.}}{12H_3PO_4} + \underset{\text{Nitric oxide.}}{20NO.}$$

Phosphoric acid is reduced to a syrupy consistence by evaporation which is the most convenient form for general use, although the Pharmacopœia (1870) recognizes only a very dilute acid. It may be obtained in the form of crystals by exposing a concentrated solution to a low temperature.

Diluted Phosphoric Acid, U. S. P.—This has the specific gravity 1.056, is colorless, inodorous, and one hundred grains of it are saturated by twenty-three and four-tenths grains of bicarbonate of potassium, the solution remaining clear. Diluted phosphoric acid has a pleasant acid taste and a strong acid reaction. It should

mix in all proportions with solution of ferric chloride without precipitation.

Impurities.—The most objectionable impurities likely to be present are the acids of arsenic, (from the phosphorus or the acids employed). *Sulphuric acid* is often present when the acid is made from bone ash; *metals*, from carelessness in preparation; and *metaphosphoric* acid, if the acid be made from glacial phosphoric acid. In this latter case sodium phosphate is also likely to be a contamination, as glacial phosphoric acid is often mixed with this substance.

Tests.—Pass sulphurous acid (sulphur dioxide) through the warm dilute phosphoric acid, then boil the solution; this will reduce the arsenic acid (if present), forming arsenious acid. Then pass sulphide of hydrogen into the liquid and a yellow precipitate will fall (trisulphide of arsenic); filter, and the solution will be *free* from *arsenic*. Add solution of albumen to a little diluted phosphoric acid, a gelatinous precipitate indicates *meta-phosphoric* acid. This impurity is also shown if a gelatinous precipitate results from mixing dilute phosphoric acid with officinal tincture of chloride of iron. Mix equal bulks of dilute phosphoric acid with dilute nitric acid, and add a few drops of solution of barium nitrate, a white precipitate shows the presence of *sulphuric acid*. In the absence of arsenious acid, if a few drops of solution of permanganate of potassium be added to dilute phosphoric acid and the mixture be boiled, it will decolorize if *phosphorous acid* be present. Phosphorous acid, if present, will also render dilute phosphoric acid turbid when boiled with a few drops of corrosive sublimate solution.—(Hoffman). Pass sulphide of hydrogen into dilute phosphoric acid and a yellow precipitate will indicate *arsenious acid;* a brown or black precipitate, copper, lead or tin. Subsequent filtration will separate these elements. Phosphoric acid is tribasic (H_3PO_4) and forms three classes of salts, one, two or three atoms of hydrogen being displaced by metals. As examples we may name trisodium phosphate (normal phosphate of sodium), $Na_3PO_4+2H_2O$.

Disodium phosphate† (hydrogen disodium phosphate), $HNa_2PO_4+12H_2O$. Monosodium phosphate (Dihydrogen sodium phosphate), $H_2NaPO_4+H_2O$.

Meta-phosphoric acid HPO_3 has the composition of ordinary phosphoric acid deprived of a molecule of water. It may be made by heating crystallized ortho-phosphoric acid to redness, thus:—

$$\underset{\text{Ortho-phosphoric acid.}}{H_3PO_4} = \underset{\text{Meta-phosphoric acid.}}{HPO_3} + \underset{\text{Water.}}{H_2O}.$$

Solution of meta-phosphoric acid is colorless, sour and precipitates solution of albumen, chloride of calcium and ferric chloride. The salts are called *meta phosphates*, but are not used in medicine.

Pyro-phosphoric Acid $H_4P_2O_7$.—This acid may be made by heating crystallized ortho-phosphoric acid very carefully to the temperature of 213° C. (415.4° F.) at which point two molecules of the acid give off one molecule of water, thus:—

$$\underset{\text{Ortho-phosphoric acid.}}{2H_3PO_4} - \underset{\text{Water.}}{H_2O} = \underset{\text{Pyro-phosphoric acid.}}{H_4P_2O_7}.$$

It may be produced dilute by heating phosphate of sodium to redness, which forms pyro-phosphate of sodium

$$\underset{\text{Phosphate of sodium.}}{2HNa_2PO_4} = \underset{\text{Water.}}{H_2O} + \underset{\text{Pyro-phosphate of sodium.}}{Na_4P_2O_7},$$

then dissolving it in water and precipitating the solution with lead acetate. The precipitate of pyro-phosphate of lead is to be washed well with water, and decomposed with sulphide of hydrogen, and the solution filtered to free it from the sulphide of lead. Pyro-phosphoric acid is colorless, inodorous and sour. It precipitates solution of nitrate of silver, while ortho-phosphoric acid does not, and does not act upon solution of albumen or solution of chloride of calcium (meta-phosphoric acid precipitates both). The salts of this acid are called *pyro-phosphates*, one, the pyro-phosphate of iron, rendered

† This is the ordinary phosphate of sodium of commerce, and officinal.

soluble by means of citrate of ammonium, being used in medicine.

Hypophosphorous Acid.—Formula, $H.H_2PO_2$. Molecular weight, 65.88.

This acid is made by boiling milk of lime and phosphorus together; phosphoretted hydrogen gas escapes, and hyphophosphite of calcium is formed.

$$\underset{\text{Phosphorus.}}{P_8} + \underset{\text{Water}}{6H_2O} + \underset{\text{Hydroxide of calcium.}}{3Ca(OH)_2} = \underset{\text{Hypophosphite of calcium.}}{3Ca2P(OH)_2} + \underset{\text{Phosphoretted hydrogen.}}{2PH_3}.$$

Filtration will now separate the soluble hypophosphite of calcium from the free lime, and by exactly neutralizing the solution with oxalic acid hypophosphorous acid is liberated. This is now to be again filtered from the insoluble oxalate of calcium and the solution of hypophosphorous acid concentrated by gentle evaporation. In this manner it may be reduced to a syrupy consistence. Hypophosphorous acid may be made according to Prof. Proctor's formula by dissolving 480 grains of calcium hypophosphite in 6 ounces of water, and mixing the solution with a solution of 350 grains of oxalic acid in three ounces of water, filtering, washing the magma of oxalate of calcium (see page 70, for direction for washing precipitates) with a little water, and afterward evaporating the mixed filtrates to the bulk of eight and one-half fluid ounces. This will make a 10 per cent. solution of hypophosphorous acid. Hypophosphorous acid is seldom or never obtained in a pure state, owing to the fact that it readily absorbs oxygen. It is recorded that explosions have been known when concentrating this acid by evaporation. The hypophosphorous acid of commerce is dilute, the usual strength being 10 per cent. It has a sour pleasant taste, is inodorous and forms salts known as *hypophosphites*. When hypophosphorous acid is strongly heated it decomposes into phosphoretted hydrogen and phosphoric acid, thus:

$$\underset{\text{Hypophosphorous acid.}}{2H_3PO_2} = \underset{\text{Phosphoretted Hydrogen.}}{PH_3} + \underset{\text{Phosphoric acid.}}{H_3PO_4}.$$

The solution of both hypophosphorous acid and the salts of hypophosphorous acid reduces solutions of the salts of gold and silver, the metals precipitating and phosphoric acid remaining.

Phosphorous Acid (H_2HPO_3) is not used in medicine. It is made by burning phosphorus in a limited supply of dry air whereby *phosphorus trioxide* (P_2O_3) is produced, and this when thrown into water combines with it, forming phosphorous acid, thus:—

$$\underset{\text{Phosphorus trioxide.}}{P_2O_3} + \underset{\text{Water.}}{3H_2O} = \underset{\text{Phosphorous acid}}{2H_3PO_3}.$$

It may be also made by the decomposition of trichloride of phosphorus by means of water. Phosphorous acid is of interest to pharmacists and physicians, since it may be present in diluted phosphoric acid which was made with an insufficient amount of nitric acid. (See detection p. 111.) Phosphorous acid forms salts which are known as *phosphites.*

Pyrogallic Acid.—Formula, $C_6H_6O_3$. Molecular weight, 125.70.

This substance is prepared by heating gallic acid, whereby at the temperature of about 210° C. a sublimate appears; this is *pyrogallic acid.* It is of a white feathery appearance and has feeble acid properties. Pyrogallic acid does not precipitate solution of gelatin. It dissolves in alcohol, ether or water, but is insoluble in chloroform; it has an acrid bitter taste and forms a blue solution with ferrous sulphate (green vitriol), and a red solution with ferric chloride. Pyrogallic acid is only occasionally used in medicine, but enters into the composition of many hair dyes.

Pyroligneous Acid.

This is not a definite acid, but is a mixture of substances obtained by the destructive distillation of wood within close retorts. It is composed of creosote, acetic acid and other products of the decomposition, in varying proportions, forming a dark colored liquid which rises

upon the tarry products of the distillation. This liquid (so-called pyroligneous acid), has a strong penetrating smoke-like odor and an acrid disagreeable taste. It is a powerful antiseptic, and if a small proportion be added to water it will prevent the decomposition of flesh that is dipped into the solution. It is occasionally used in medicine diluted, as a wash, but the chief use of it at present is the production of acetic acid.

Salicylic Acid.—Formula, $H_2C_7H_4O_3$. Molecular weight, 137.67.

This substance derives its name from salicin, from which it was originally prepared. It is made in large amounts by a patented process, whereby solution of caustic soda is mixed with carbolic acid and the mixture evaporated to dryness. This is then heated to 100° C. (212° F.), and while a constant stream of carbon dioxide passes through the mixture the temperature is increased until it reaches 250° C. (482° F.). Much carbolic acid distills and the contents of the retort after the reaction, consists of basic salicylate of sodium, which may be decomposed by means of hydrochloric acid, and the salicylic acid afterward purified. Wintergreen oil consists mainly of methyl-salicylic acid, or at least it may be easily separated into methyl-alcohol and salicylic acid, the result being a very pure acid.

Preparation.—Dissolve three parts of white caustic potash in two parts of water in a glass or porcelain vessel; heat to 82° C. (179.6° F.), and stir into it three parts of wintergreen oil (pure), then pour the mixture into sixty-four parts of cold distilled water previously mixed with eight parts of hydrochloric acid, (all proportions by weight). A mass of silky white crystals of salicylic acid will separate, which needs only to be drained, dissolved in alcohol and recrystallized. Salicylic acid scarcely dissolves in cold water, but is soluble to a considerable extent in warm water. It dissolves freely, and without residue, in alcohol and in ether. Borax increases its solubility in water, but the solutions soon decompose. Ferric chloride is a delicate test for salicylic acid, producing a

deep red or violet color as the solution of salicylic acid is concentrated or dilute, and even very dilute solutions turn distinctly pink. Either acids or alkalies destroy this color. The salts of salicylic acid are known as *salicylates*, the sodium salicylate being used in medicine. Neither salicylic acid nor salicylate of sodium should possess the odor of carbolic acid.

Silicic Acid.—Formula, H_4SiO_4 or $Si(OH)_4$. Molecular weight, 95.84.

This acid is not known in a pure state. It corresponds with a molecule of silicon dioxide (SiO_2), united with two molecules of water, thus:—

$$\underset{\text{Silicon dioxide.}}{SiO_2} + \underset{\text{Water.}}{2H_2O} = \underset{\text{Silicic acid.}}{H_4SiO_4}.$$

As with sulphurous acid (H_2SO_3) and carbonic acid, (H_2CO_3), both of which have a strong tendency to split up into water and the oxide, so it is with silicic acid. Thus we find that even slight evaporation will decompose all three as explained by the following equations:—

$$\underset{\text{Sulphurous acid.}}{H_2SO_3} = \underset{\text{Water.}}{H_2O} + \underset{\text{Sulphur dioxide.}}{SO_2}.$$

$$\underset{\text{Carbonic acid.}}{H_2CO_3} = \underset{\text{Water.}}{H_2O} + \underset{\text{Carbon dioxide.}}{CO_2}.$$

$$\underset{\text{Silicic acid.}}{H_4SiO_4} = \underset{\text{Water.}}{2H_2O} + \underset{\text{Silicon dioxide.}}{SiO_2}.$$

Preparation.—To a dilute solution of silicate of sodium (soluble glass) add an excess of hydrochloric acid. This will produce chloride of sodium and silicic acid. By dialysis the excess of hydrochloric acid and the chloride of sodium may be separated, thus giving us a clear solution of silicic acid.

Properties.—Solution of silicic acid prepared as above has a slight acid reaction, but is almost or quite tasteless. It can be somewhat concentrated by boiling in a close vessel, but soon decomposes. A five per cent. solution gelatinizes in a few days. When dried over sulphuric acid a

glass-like substance remains of the composition H_2SiO_3, which has been named *meta-silicic* acid. Silicic acid is not used in medicine, but a salt of it, silicate of sodium (soluble glass) is sometimes employed in surgery or chemical experiments. The salts of silicic acid are known as **silicates.**

Succinic Acid.—Formula, $H_2C_4H_4O_4$. Molecular weight, 117.72.

Succinic acid is produced by the distillation of amber, but it is found in several plants and occasionally in animals. It may be made by fermentation of malic acid as follows: Crude malate of calcium is mixed with water, and yeast, or old cheese, when succinate of calcium forms, which by decomposition with dilute sulphuric acid yields succinic acid and sulphate of calcium. To prepare it from amber place coarsely powdered amber in a glass flask which is closely connected with a large cooled receiver. The flask must be entirely surrounded with sand. Apply heat by means of a sand-bath and increase it to 500° F. and continue it as long as vapors of acid are evolved. The product will be oil of amber, acetic and succinic acids. The most of the succinic acid condenses in the neck of the flask, and it may be purified by repeated solutions in water, filtration, and crystallization.

Properties.—Succinic acid is white, odorless, slightly soluble in cold water and very soluble in boiling water. As found in commerce it is often of a dirty brown color and has a disagreeable odor of amber oil, being totally unfit for use. Succinic acid forms soluble salts with the alkalies; these are known as *succinates.*

Sulphocarbolic Acid.—Formula, $HC_6H_5SO_4$. Molecular weight, 173.64.

Synonyms:—Phenol-sulphonic acid. Phenyl-sulphuric acid.

This acid is made by mixing equal weights of sulphuric and carbolic acids and heating to 50° C. (122° F.) This produces what is known as *ortho-sulphocarbolic acid.*

If the heat be increased to 149° C. (300° F.) a molecular change takes place, *para-sulphocarbolic acid* being formed. These acids are not known in an absolute state, but are always contaminated with free sulphuric acid and carbolic acid. Sulphocarbolic acid forms salts known as *sulphocarbolates*, and by the crystallization of such the acid may be obtained united with a base. It decomposes, however, when liberated from such combinations.

Sulphuric Acid.—Formula, H_2SO_4. Molecular weight, 97 82.

Synonym.—Oil of vitriol.

This acid is a compound of sulphur trioxide SO_3 (sulphur peroxide) and water, thus:—

$$\underset{\text{Sulphur trioxide.}}{SO_3} + \underset{\text{Water.}}{H_2O} = \underset{\text{Sulphuric acid.}}{H_2SO_4}.$$

Sulphuric acid is manufactured in enormous quantities in all civilized countries, and is consumed in almost every branch of industry. It has been known since a very early day, deriving the name *oil* of vitriol from the fact that at first it was produced as an oily liquid, by distilling vitriol (sulphate of iron). Sulphuric acid is practically made by passing sulphur dioxide into a leaden chamber, in connection with steam, air, and nitrous fumes. The reaction may be explained as follows:—

$$\underset{\text{Nitrogen peroxide.}}{NO_2} + \underset{\text{Sulphur dioxide.}}{SO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Sulphuric acid.}}{H_2SO_4} + \underset{\text{Nitric oxide.}}{NO}$$

The nitric oxide (NO) thus liberated seizes upon free oxygen of the atmosphere, and forms nitrogen peroxide (NO_2) again, thus:—

$$\underset{\text{Nitric oxide.}}{2NO} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Nitrogen peroxide.}}{2NO_2}.$$

Theoretically a limited amount of nitrogen peroxide should produce an indefinite amount of sulphuric acid, as it (NO_2) simply acts as a carrier of oxygen from the atmosphere to the sulphur dioxide. In practice, however, this is not the case, for the air admitted to the chamber

contains four-fifths of its bulk of *nitrogen*, and as the nitrogen escapes from the chamber to make room for fresh supplies of air, a portion of all the other gases is lost with it. Thus we find that in the practical manufacture of sulphuric acid, the vapors of burning sulphur, nitrous fumes, steam, and atmospheric air are constantly passing into the leaden chamber.† This results in an aqueous solution of sulphuric acid and after some time the liquid is drawn off and evaporated to this proper consistency in glass or platinum.

Properties.—Sulphuric acid of commerce is colorless, odorless, and as prepared in accordance with the U. S. P. of the specific gravity of 1.843. It contains 79 per cent. of sulphur trioxide (SO_3) and 96.8 per cent. of pure sulphuric acid, (H_2SO_4). It is one of the strongest of acids, especially attacking those organic bodies that contain water, or the elements of water (oxygen and hydrogen,) and decomposing them with ease. Thus, wood is at once blackened by it from liberation of carbon, and flesh is destroyed by actual disintegration. Sulphuric acid unites with water, the mixture contracts and much heat is liberated. The salts of sulphuric acid are called *sulphates*.

Diluted Sulphuric Acid, U. S. P., is made by gradually mixing two troy ounces of sulphuric acid with enough distilled water to make the mixture measure a pint, then after twenty-four hours filtering through paper. The specific gravity will be 1.082.

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† Some prefer to explain the reaction as follows: The steam and sulphur dioxide unite to form sulphurous acid thus:

$$SO_2 + H_2O = H_2SO_3,$$

while the action of sulphuric acid on nitrate of potassium produces nitric acid, thus:

$$2KNO_3 + H_2SO_4 = K_2SO_4 + 2HNO_3.$$

These substances pass together into the chamber and the reaction which follows may be explained as follows:

$$3H_2SO_3 + 2HNO_3 = 3H_2SO_4 + H_2O + (NO)_2.$$

Undoubtedly both nitric acid and nitrogen peroxide are evolved from the saltpetre, and therefore, we may possibly accept both of the explanations, but there is every reason to believe the reaction is much more complex.

Pure Sulphuric Acid (H_2SO_4) can not be made by evaporation or distillation, for even if officinal sulphuric acid is distilled, the last third of the distillate will contain two per cent. of water. However, if this is cooled, crystals of pure sulphuric acid separate, of the specific gravity 1.854 at 0° C. (32° F.)

Fuming Sulphuric Acid.—This is known as *Nordhausen sulphuric acid*, and was made before sulphuric acid was produced from sulphur. It derives its name from Nordhausen where it was first prepared, being made by roasting green vitriol and then distilling. It is in reality a solution of *sulphur trioxide* (SO_3, anhydrous sulphuric acid) in sulphuric acid. It is generally of a brown color, of the specific gravity 1.86 to 1.89 (R. & S.), and gives off sulphur trioxide (SO_3) which immediately unites with the vapor of the atmosphere to form sulphuric acid, dense white fumes exhibiting themselves, thus:—

$$\underset{\text{Sulphur trioxide.}}{SO_3} + \underset{\text{Water.}}{H_2O} = \underset{\text{Sulphuric acid}}{H_2SO_4}.$$

From this reason the acid is known as *fuming sulphuric acid.* Some prefer to call this acid *disulphuric acid* $H_2S_2O_7$, and indeed salts of this composition can be made, as for example disulphate of sodium $Na_2S_2O_7$.

Tests and Impurities.—Sulphuric acid, or the solution of a soluble sulphate, will throw down a white precipitate (sulphate of barium) upon the addition of a solution of chloride of barium, which precipitate will *not* dissolve in nitric or in hydrochloric acid. (Sulphite of barium dissolves upon the addition of these reagents.)

Lead is sometimes present in sulphuric acid. It may be shown by mixing one part of sulphuric acid slowly with four or five parts (all by measure) of alcohol, when, if lead is present, a white turbidity, or precipitate, will follow.

Arsenic, often an impurity of sulphuric acid, may be detected by Marsh's test. *Sulphurous acid* will decolorize a strip of blue iodide of starch paper, if suspended above the suspected acid in a close bottle; and *oxides of nitrogen* will cause, under like conditions, a blue colora-

tion to appear upon paper which has been moistened with iodide of potassium and starch solution.

Sulphuric acid, when much diluted, has an agreeably acid taste, and is occasionally used in hot weather as a cooling drink. Concentrated sulphuric acid is a powerful poison. It decomposes the membranes of the throat and stomach, and may even destroy life before reaching the stomach. It is the most fearful of the corrosive mineral acids, and if a large dose has been swallowed, the effects can not be antidoted, for it actually disintegrates the vital organs with which it comes into contact.

Treatment.—Chalk made into a milk with water should be freely administered, or mixtures of carbonate or calcined magnesia with water. Solution of carbonate of sodium or bicarbonate of sodium in milk or water is of value, and such drinks as gruel, and linseed oil mixed with lime water, milk, etc., to mechanically envelop the poison. For reasons before stated, little hope of recovery need be entertained if a large dose has been swallowed.

Sulphurous Acid.—Formula, H_2SO_3. Molecular weight 81.86.

Sulphurous acid is made by passing sulphur dioxide into water, thus:

$$\underset{\text{Sulphur dioxide.}}{SO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Sulphurous acid.}}{H_2SO_3}$$

Sulphurous acid is not known in a pure state, for when its solution in water is evaporated, the acid decomposes into water and dioxide of sulphur.

Preparation.—Heat in a glass retort a mixture of sulphuric acid eight parts, and powdered charcoal one part, and pass the gas into *cold* water. The reaction is shown by the following equation:—

$$\underset{\text{Sulphuric acid.}}{4H_2SO_4} + \underset{\text{Carbon.}}{C_2} = \underset{\text{Carbon dioxide.}}{2CO_2} + \underset{\text{Water.}}{4H_2O} + \underset{\text{Sulphur dioxide.}}{4SO_2}.$$

The sulphur dioxide is absorbed by the water to form sulphurous acid, in the manner previously explained.

Properties.—Solution of sulphurous acid is colorless; its vapor is irritating and suffocating when inhaled. It is

not corrosive, evaporates readily, and for this latter reason must be securely stopped. It also absorbs oxygen with avidity, which is another inducement for excluding the atmosphere, as it passes, in this case, into sulphuric acid, thus:—

$$\underset{\text{Sulphurous acid.}}{2H_2SO_3} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Sulphuric acid.}}{2H_2SO_4}.$$

Sulphurous acid forms salts known as *sulphites*, which evolve sulphurous acid upon addition of hydrochloric acid, or dilute sulphuric acid, and which may then be recognized by its odor. To a solution of a neutral soluble sulphite add a solution of chloride of barium, and a white precipitate will fall, which will *dissolve* upon the addition of hydrochloric acid. (Sulphate of barium will not dissolve.)

Tannic Acid.—Formula, $C_{14}H_{10}O_9$. Molecular weight, 321.22.

This is obtained from nutgalls by extracting the powdered galls with ether, which has been agitated with water, and subsequently evaporating the ether. The tannin will remain as a light porous amorphous substance of a greenish or brownish color. Tannic acid from nutgalls is freely soluble in water, alcohol, glycerin, or in ether which is mixed with alcohol or which contains water, (washed ether). It is very astringent and the aqueous solution has an acid reaction. It precipitates most alkaloids as insoluble, or almost insoluble tannates; also in like manner the salts of most metals. Solutions containing gelatin are precipitated by tannin, consequently gelatin is an excellent agent to separate tannin from aqueous solutions, and in most cases a shred of gelatin or rawhide will remove it in a few hours. Tannin of commerce has a peculiar odor, which, according to Prof. Proctor, may be removed by treatment with benzine. Tannic acid precipitates solutions of ferric salts, *blue-black*. Tannic acid changes under the action of continued moisture into *gallic* acid (see gallic acid), and hence is now sometimes called digallic acid.

Tannins.—There are a multitude of tannins. The oak

bark tans leather, and produces blue-black precipitates with ferric salts, but (distinction from tannin) will not yield pyrogallic acid. Most plants contain a tannin in some part of their organism, and but few of these principles agree in all characters. Some plants contain two distinct tannins. As a means of classification it is customary to speak of tannins as yielding blue or green precipitates with ferric or with ferrous salts; even this is indefinite and uncertain, for a variation in the proportion of the tannin, or even the influence of the acid with which the iron is united, will modify the reaction or render it perfectly unreliable. Some tannins are comparatively permanent, others, like tannin from nutgalls, are easily decomposed. Of the chemical constitution of the tannins we know but little.

Tartaric Acid.—Formula, $H_2C_4H_4O_6$. Molecular weight, 149.64.

Tartaric acid exists in many fruits. It is a constituent of grapes, and in small amounts is found in many roots, barks and herbs, where it exists either free or as the acid tartrate of potassium, $KHC_4H_4O_6$. When the juice of grapes is permitted to ferment the acid tartrate of potassium separates in consequence of its insolubility in an alcoholic menstruum, and adheres to the sides of the cask in the form of an incrustation, (known as argol or crude tartar). This is dissolved and then decomposed by means of carbonate of calcium, and the tartrate of calcium thus formed is in turn decomposed with sulphuric acid. This gives rise to sulphate of calcium and solution of tartaric acid, which is separated from the sulphate of calcium by filtration. The solution of tartaric acid is afterward evaporated until the tartaric acid crystallizes.

Properties.—Tartaric acid will dissolve in its weight of cold water and in half its weight of boiling water; alcohol dissolves one-third its weight. Its crystals are colorless and odorless, have a sharp sour taste and do not contain water of crystallization. It unites with bases to form both acid and neutral salts known as *tartrates*.

Impurities and Tests —Inorganic salts may be shown by dissolving the acid in six parts of alcohol; admixtures will remain *undissolved*.. If it contain *calcium* salts a dilute solution of the acid previously neutralized with ammonia will precipitate *white* with solution of oxalate of ammonium.

Sulphates are shown by a white precipitate when a dilute solution of tartaric acid is mixed with a little solution of nitrate of barium. Metallic impurities, such as copper or lead, are thrown down when its aqueous solution is treated with sulphide of hydrogen.

Valerianic Acid.—Formula, $HC_5H_9O_2$. Molecular weight, 101.77.

Valerianic acid (valeric acid) was first found in valerian root, and from thence derived its name. It has also been obtained from other plants, and at present is artificially prepared in quantities (none of the true is in market), being produced from amyl alcohol; thus with caustic potash the process is explained by the following equation.

$$\underset{\text{Amyl alcohol.}}{C_5H_{12}O} + \underset{\text{Caustic potassium.}}{KOH} = \underset{\text{Valerianate of potassium.}}{KC_5H_9O_2} + \underset{\text{Hydrogen.}}{2H_2}.$$

It is usually made from a mixture of dichromate of potassium, sulphuric acid, amyl alcohol (fusel oil), and water. The reaction is violent, however, without extra precautions, and as the odor of this acid is very offensive, few care to experiment with it. The modern name for valerianic acid is *isopentoic* acid, or *isopropyl-acetic* acid.

Properties.—Valerianic acid is thin, colorless, oily, and offensively disagreeable in odor to most persons, either in large or small amount. It dissolves freely in alcohol, and this solution is almost odorless. It is also very soluble in water; and a few grains will give the odor to a considerable amount of water. Valerianic acid differs considerably in specific gravity and other properties, in accordance with its origin. The salts of valerianic acid are called *valerianates*.

Impurities and Tests.—This acid must dissolve in from 26 to 30 parts of water. If it dissolves in a less amount

than 26 parts, it is contaminated either with alcohol, acetic acid, or butyric acid. (Hoffmann.) If it requires more than 30 parts, the acid contains impurities less soluble in water than valerianic acid. Neutralize a solution of valerianic acid with an alkali, shake the mixture with ether, filter, and allow the ether to evaporate spontaneously, and amyl alcohol, valeric aldehyde, or neutral ethers will, according to Allen, remain after the ether has escaped.

CHLORINE.

Symbol Cl. Atomic weight, 35.37.

Scheele discovered chlorine in 1774, and named it "*dephlogisticated marine acid gas,*" from the fact that he prepared it from *marine acid* (hydrochloric acid). Berthollet afterward endeavored to prove that it was a compound of hydrochloric acid gas and oxygen, but Davy (1810) disproved his assertion. Chlorine derived its name from its greenish yellow color (named by Davy). It is always found in combination in nature, but in this form exists in immense amounts, notably combined with potassium, sodium or magnesium. Sodium chloride constitutes the largest part of the solid constituents of the waters of the ocean, of the waters of many inland lakes, and is found in brine springs, and as rock salt is found in immense beds in many countries. The chlorides of the alkaline and the alkali earth metals are found in both animal and vegetable organisms and seem necessary to health and existence. Chlorine gas is most easily prepared by warming hydrochloric acid with manganese dioxide, the operation being explained as follows :—

$$\underset{\text{Hydrochloric acid.}}{4HCl} + \underset{\text{Manganese dioxide.}}{MnO_2} = \underset{\text{Chlorine gas.}}{Cl_2} + \underset{\text{Manganous chloride.}}{MnCl_2} + \underset{\text{Water.}}{2H_2O}.$$

Experiment.—Place one part of powdered manganese dioxide in a chemical flask capable of holding sixteen parts, and pour upon it six parts of hydrochloric acid previously diluted with half its weight of water. Connect the flask by means of a rubber stopper perforated

and fitted with a glass tube, with two wash bottles, the first of which contains two parts of distilled water, the second, four parts of sulphuric acid. From the second wash bottle a tube passes to the bottom of a bottle filled with air. Heat must now be gently applied to the flask, by means of an alcohol lamp or gas jet, and a steady flow of chlorine will follow. When the air in the receiving bottle is displaced by chlorine, the heat must be removed from the generator. If chlorine be inhaled, the result is a distressing irritation, therefore, it is advisable that the experiment be conducted beneath a good flue, or in the open air, and the connections must be secure and tight. If the chlorine be conducted into cold distilled water, it is dissolved to a great extent, and when the water is saturated with the gas the result is *aqua chlorinii*, or chlorine water of the Pharmacopœia.

Chlorine has the atomic weight, 35.37. It is a transparent greenish-yellow gas of a disagreeable suffocating odor. It liquefies under the pressure of six atmospheres at 0° C. (32° F.) or at the ordinary pressure of the atmosphere when it is reduced to —34° C. (—29.2° F.) Chlorine is not inflammable and does not unite directly with oxygen. It unites directly with hydrogen, hydrochloric acid resulting, thus:—

$$Cl_2 + H_2 = 2HCl.$$

This affinity for hydrogen explains the bleaching power of chlorine, for the chlorine decomposes the water present, unites with the hydrogen, and the oxygen which is liberated in a nascent state, oxidizes and destroys the coloring matters. If a jet of burning hydrogen be thrust into a jar of chlorine gas, the combustion continues attended by the evolution of both light and heat; the product is hydrochloric acid. If finely powdered metallic antimony be sprinkled into a jar of chlorine gas, the particles form a shower of sparks, the two elements uniting to produce chloride of antimony. A piece of filtering paper saturated with oil of turpentine bursts into a flame when it is plunged into a jar of chlorine gas. The chlorine decomposes the oil, unites with the hydrogen and liberates the

carbon. Chlorine unites with many metals to form *chlorides.* A definite hydrate, $Cl.5H_2O$, is produced when chlorine gas is passed to saturation into water just above 0° C. (32° F.) This is a solid crystalline compound and is known as chlorine hydrate. The solution of chlorine in water has a greenish color and smells strongly of chlorine. A portion of the water decomposes by age, the chlorine uniting with its hydrogen to form hydrochloric acid, with liberation of oxygen, thus:—

$$\underset{\text{Chlorine.}}{2Cl_2} + \underset{\text{Water.}}{2H_2O} = \underset{\text{Hydrochloric acid.}}{4HCl} + \underset{\text{Oxygen.}}{O_2}.$$

There are three compounds of oxygen and chlorine, all made by indirect processes, as the two elements refuse to unite directly. They are:—

Cl_2O, chlorine monoxide, or hypochlorous oxide.
Cl_2O_3, chlorine trioxide.
Cl_2O_4 or ClO_2, chlorine tetroxide.

There are four compounds of chlorine with oxygen and hydrogen, known as the *oxy acids* of chlorine, they are:—

$HClO$ hypochlorous acid.	$HClO_2$ chlorous acid.
$HClO_3$ chloric acid.	$HClO_4$ perchloric acid.

Chlorine combines directly with carbon and other elements in certain organic compounds, for example chloroform, $CHCl_3$, and such do not precipitate a solution of silver nitrate. We have considered the compound formed of chlorine and hydrogen (hydrochloric acid), under the acids, and shall study the *chlorides* which are used in medicine, under the respective electro-positive element with which chlorine is united to form each compound. In the section of our work devoted to *organic chemistry,* chloral hydrate, chloroform, etc., will be considered.

Tests for free Chlorine.—If a solution containing free chlorine be mixed with a solution of ferrous sulphate and potassium sulphocyanide, a red color results. If a solution of chlorine, free from nitrous acid, be mixed with a solution of iodide of potassium and starch paste, a blue color at once results. Ozone also gives this reaction.

Tests for Soluble Chlorides.—Mix solution of nitrate of silver with the solution of any chloride, (chloride of sodi-

um for example); a white precipitate results. Decant the supernatant liquid, add excess of strong nitric acid and boil; the precipitate remains. Decant the acid, add ammonia water in excess to the precipitate, and solution results.

Tests for insoluble Chlorides.—Boil calomel, for example, with solution of caustic potash; permit the result to settle; decant the clear solution and then add to it nitric acid until in slight excess. This will give a solution of chloride of potassium which may be detected by the foregoing test for soluble chlorides.

BROMINE.

Symbol, Br. Atomic weight, 79.75.

Bromine was discovered by Balard, in 1826, who obtained it from the mother liquor of sea-water from which salt had been separated. Bromine does not occur free in nature, usually being combined with the alkaline metals, or the metals of the alkaline earths, forming *bromides*. The source of supply was formerly certain salt springs of Europe, and the potash beds of Stassfurth. Now the United States is supplied from the mother liquors of the salt springs of the Kanawha river, in this country. The preparation of bromine is very simple: the mother liquor of the salt springs is mixed with sulphuric acid and black oxide of manganese, in stone retorts, and heat is applied. This liberates chlorine, which reacts upon the soluble bromides present, decomposes them, with liberation of bromine and production of chlorides. The bromine distills, and is condensed. Bromine is liquid at ordinary temperatures, mercury and bromine being the only elements of which this can be said. It is black in bulk, but when shaken against the sides of the bottle, shows in thin layer a deep ruby red color. It is three times as heavy as water, having a specific gravity 3.1872. At —22° C. (—7.6° F.) it solidifies. It is volatile, evolving heavy red fumes when exposed to the air, and these are very irritating to the eyes, throat, and lungs. The best antidote for the effects of inhaled bromine is to breathe ammo-

nia vapor, and it is customary for workmen to keep a vessel of ammonia water convenient while working with bromine. Bromine produces a painful sore, if it be permitted to come into contact with the skin. In such cases the part attacked should at once be washed with dilute ammonia water, and then treated with oil or bland cerates. A small amount of the vapor of bromine in the air reminds one by its odor of chlorine, and it bears a close resemblance to that element in other respects. Bromine bleaches organic bodies, but less energetically than chlorine. It combines directly with potassium with violence, but does not affect metallic sodium at ordinary temperatures.

When bromine is dissolved in water it imparts an orange-red color, and when a saturated solution is made at just above 0° C. (32° F.) a definite crystalline body results, of the composition $Br5H_2O$. Bromine water decomposes readily if exposed to sun-light, hydrobromic acid being formed, and oxygen being liberated. Bromine and hydrogen unite to form hydrobromic acid (HBr), but this compound is treated of under *acids*, to which the reader is referred.

Bromine and Chlorine, or *Chloride of Bromine*, has been used by physicians. Under this name a mixed solution of the two substances is prepared. The composition, however, of "bromide of chlorine" is in doubt. The solution is made by mixing one part of bromine with sixty parts of cold water, and passing chlorine gas into it until the bromine disappears and a dark-red solution is formed. Chloride of bromine has an intensely irritating odor, and is said to be a powerful caustic. The compounds of bromine with single elements are known as *bromides*. There is no known oxide of bromine.

Tests for Soluble Bromides and Bromine.—If a solution of nitrate of silver be mixed with a solution of a bromide, a yellowish precipitate results. Decant the solution, add dilute nitric acid and warm; the precipitate remains, and ammonia water in slight excess partially dissolves it. Iodides give the same reaction. To the solution of a

bromide add chlorine water cautiously, care being taken to avoid an excess. Bromine is liberated and chlorine takes its place, a red, or yellowish-red solution resulting. Add a small amount of chloroform, or carbon disulphide, and agitate; the bromine is dissolved and the layer of lower solution will contain it, being colored red or yellowish in proportion to the bromine present. Decant the overlying aqueous solution, suspend a slip of moistened starch paper in the test-tube, and warm gently; the paper becomes yellow, which color disappears when the paper is exposed to the air.

IODINE.

Symbol I. Atomic weight, 126.53.

In the year 1812, M. Courtois, of Paris, discovered iodine in the mother liquor of the soda-salts prepared from kelp. The name was derived from the violet color of its vapor. Iodine occurs in nature combined with other elements, chiefly potassium, sodium, and calcium, forming iodides. The waters of the ocean are its great reservoir; it is obtained from the ashes of sea weeds. Sponge also yields iodine, and at one time the ashes (*burnt sponge*) were used in medicine. From the mother liquor of the lixivated ashes of sea weeds, iodine is obtained by processes similar to that mentionedfor making bromine.

Iodine is found in commerce as bright, black or bluish-gray plates, and is sold under the name *re-sublimed iodine*. It has the specific gravity 4.948, melts at from 113° to 115° C. (235.4° to 239° F.) When heated it boils at 200° C. (392° F) giving rise to violet vapors, and from this fact, as we have said, the element derived its name. Iodine is a bad conductor of electricity. It dissolves sparingly in water, but freely in solution of iodides and in alcohol, yielding brownish-red solutions. Iodine acts energetically upon many volatile oils. With oil of juniper the compound has a greenish or yellowish color, being known as *iodized oil of juniper*. The chemical relations of iodine are with chlorine and bromine. If iodine imparts a red or a deep yellow color to pure water, the

presence of chlorine may be inferred. The important characteristic of iodine is the power it possesses of striking a deep blue color with starch paste. This experiment is illustrated by boiling a few grains of starch with a little water, and adding a drop of tincture of iodine to the solution. In this connection it must be remembered that the iodine *must* be free and the starch granules burst and in solution. This fact may be illustrated by adding a few grains of iodide of potassium to a little starch paste; there will be no coloration, but the addition now of a little chlorine water at once produces a deep blue color by liberation of iodine.

The compound of iodine and hydrogen, hydriodic acid (HI), has been considered under the head of acids. Chlorine forms, with iodine, two compounds, iodine monochloride (ICl), and iodine trichloride (ICl_3). Neither is used in medicine. Iodine combines with bromine, to form a compound which is used in medicine under the name—

Bromide of Iodine.—This is used in form of solution, and is made by rubbing one part of iodine into a paste with a little water, then mixing with it two parts of bromine, and when the iodine is dissolved diluting the solution to 60 parts. (Wittstein). Some prefer a more concentrated solution. Solution of bromide of iodine decomposes by exposure to light, hydrobromic and hydriodic acids being formed and oxygen evolved. The medicinal compounds of iodine with the metals are considered in their proper place. Iodine unites with oxygen to form *pentoxide* of *iodine* (I_2O_5), and this in turn combines with water to form iodic acid, thus:—

$$I_2O_5 + H_2O = 2HIO_3.$$

Tests for Soluble Iodides.—1st. Mix a solution of an iodide with a little thin starch paste, and acidulate the mixture with hydrochloric acid. Now add a little solution of *nitrite* of potassium; if a *trace* of iodine be present a blue color results in a few seconds; if the iodine is present in appreciable amount the blue color appears immediately.

2d. To a solution of one part of sulphate of copper, add a solution of two and one-half parts of ferrous sulphate. Mix this with a neutral solution of an iodide; a white precipitate of cuprous iodide (CuI) results. This serves to distinguish an iodide from chlorides and bromides, as they do not respond to the test.

Test for Free Iodine.—It will be observed that the first test results in liberation of iodine, which is detected by starch paper. In all cases where free iodine is present a blue color is produced with starch paste, or starch paper.

FLUORINE.

Symbol F. Atomic weight, 19.1.

The element which gives rise to the *fluorine* compounds, has never been isolated. Numerous attempts with the object of its identification have been made, and by the most careful experimenters, but all resulted in complete failure. Fluorine combines with hydrogen, silicon and the metals with avidity. From this reason when the element is liberated from combination it at once unites with the vessel which contains it, even if such be platinum, glass or carbon. Fluorine is remarkable for the fact that as yet it has never been made to combine with oxygen, and thus stands isolated from all other elements. There is no compound of fluorine of interest from a therapeutical point. It unites with the metals to form *fluorides* of which the fluorides of the alkaline earths are insoluble in water. Fluorides combine among themselves to form double fluorides, and as a rule the fluorides dissolve in hydrofluoric acid.

Hydrofluoric Acid (HF) is of interest since its solution in water forms the well known liquid employed to etch upon glass. It is a very dangerous substance to experiment with in a concentrated form. The inhalation of its vapor is dangerous, and it is a matter of record that Prof. Nickles, of Naples, lost his life from accidental exposure to it. Ordinary hydrofluoric acid is a solution of the pure acid in water. It is made by acting upon the mineral fluor spar (CaF_2), with sulphuric acid, and passing

the liberated hydrofluoric acid into water. For this purpose a leaden still is employed, and the solution of hydrofluoric acid is afterward retained in rubber bottles. Great care must be employed in its preparation lest the operator inhale the fumes.

Hydrofluoric acid, if anhydrous, is colorless and volatile and does not act upon glass. If moisture be present, however, the glass is corroded. It is very soluble in water. The property hydrofluoric acid possesses of etching glass may be exemplified by placing a little powdered fluor spar in a leaden cup and covering the cup with a waxed plate of glass, through the wax of which a figure is scratched with some hard blunt instrument. Pour a little sulphuric acid upon the fluor spar and in a short time the glass will be corroded where the wax is removed. The hydrofluoric acid is produced as follows:—

$$\underset{\text{Fluor spar.}}{CaF_2} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Hydroflu. oric acid.}}{2HF.} + \underset{\text{Calcium sulphate.}}{CaSO_4}.$$

The experiment must be performed under a good flue, or in the open air, as it is necessary that the operator should not inhale the vapor.

OXYGEN.

Symbol O. Atomic weight, 15.96.

Oxygen is the most abundant of the elements. It forms (in combination) 88.87 per cent. by weight of all the water of the globe, and as a gas 23 per cent. of the weight of the air. It enters into the composition of the rocks, earths and metallic oxides which compose the solid parts of the earth, and forms a large part of all organic substances both animal and vegetable. Oxygen was discovered in 1774, by Priestly, and obtained by heating red precipitate. The name is derived from two words meaning *I produce sour*. The name was given it by Lavoisier who believed that all acids contained oxygen, (see acids).

Although oxygen is free in the atmosphere to the extent of nearly one-fourth the weight of the air, it is obtained from some body which parts with it at a moderate

temperature. For this purpose the black oxide of manganese, mercuric oxide, or chlorate of potassium, may be employed. In experiments it is very easily made as follows (fig 49):

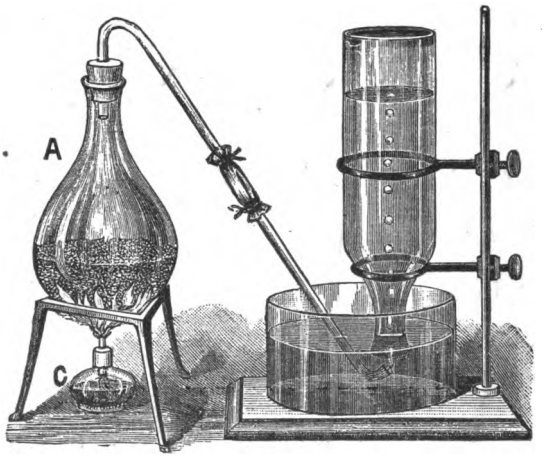

Fig. 49.

Mix two parts of powdered black oxide of manganese with three parts of powdered potassium chlorate, and place a portion of the mixture in the small chemical flask A, or even a large test tube, and connect this by means of a glass tube with the pneumatic trough B. Heat the flask with a spirit lamp, C, until the gas passes regularly, then permit it to flow into the receiving bottle. In this experiment the black oxide of manganese and the chlorate of potassium must be well dried. Half an ounce of the mixture will be sufficient to fill a quart bottle with oxygen. In this experiment the whole of the oxygen is given off from the chlorate of potassium, chloride of potassium remaining, thus:—

$$\underset{\text{Potassium chlorate.}}{2KClO_3} = \underset{\text{Potassium chloride.}}{2KCl} + \underset{\text{Oxygen.}}{3O_2}.$$

The object of the admixture of the manganese dioxide is for the purpose of liberating the oxygen at a low temperature. It has been practically demonstrated that this mixture liberates oxygen at less than the boiling point of water, while from either pure chlorate of potassium or manganese dioxide, a high heat is required. The explanation of this fact is not satisfactory, some believing that a higher oxide of manganese is constantly formed at the expense of the oxygen of the potassium chlorate, and that this is reduced in a continuous manner, while others prefer to consider it as an example of catalytic action (see catalysis, page 71).

Oxygen is colorless, odorless, and tasteless, and has the specific gravity of 1.1 when compared to air. The natural condition of the element is in the form of a gas. It was liquefied simultaneously in 1877 by Pictet and Cailletet, independently of each other. For this purpose a temperature of —130° C. and a pressure of 475 atmospheres was required. Oxygen is absorbed by many metals when in the molten state, and is given off by them upon cooling. When oxygen unites with a metal, the compound is termed an oxide. Some metals form several definite oxides.

Oxides and their Composition.

Elements differ from one another with regard to the number of oxygen atoms with which they combine. The terms which we apply to individual oxides have reference to the number of oxygen atoms in combination with an atom or atoms of the other element.

Monoxides.—These contain one atom of oxygen; thus as examples we may name—

K_2O.....Potassium monoxide.
Na_2O.....Sodium monoxide.
Li_2O.....Lithium monoxide,
BaO.....Barium monoxide.
CaO.....Calcium monoxide,
MgO.....Manganese monoxide.

Dioxides.—These contain two atoms of oxygen, and as examples we may name—

BaO_2, barium dioxide.
SrO_2, strontium dioxide.
PbO_2, lead dioxide.

Trioxides.—These contain three atoms of oxygen, thus:

Au_2O_3, gold trioxide.
Ir_2O_3, iridium trioxide.
Bi_2O_3, bismuth trioxide.

It will be noticed that each of the above-named trioxides contains two atoms of the electro-positive element united with three atoms of oxygen. To such compounds the prefix *sesqui* is often applied, meaning *one and one-half.* Thus it will be observed that one and one-half atoms of oxygen *appear* to be combined with one atom of metal in each compound, for—

$Au_2O_3 \div 2 = AuO_{1\ 1\text{-}2}$.
$Ir_2O_3 \div 2 = IrO_{1\ 1\text{-}2}$.
$Bi_2O_3 \div 2 = BiO_{1\ 1\text{-}2}$.

According to our view of the atom, this compound can not exist, as we consider the atom indivisible; therefore, the molecule of each sesquioxide consists of three atoms of oxygen united with two atoms of the other element—in reality, one and one-half times as many atoms of oxygen as there are of the electro-positive element.

Sesquioxides may then from this view be considered as intermediate between monoxides and dioxides. Formerly the term *sesqui* was universal. At present even it is extensively used; for example:—sesqui-oxide of iron, Fe_2O_3. Trioxides are also compounds in which one atom of a metal is combined with three atoms of oxygen, thus:—

WO_3, tungsten trioxide.
UO_3, uranium trioxide.
SO_3, sulphur trioxide.
MnO_3, manganese trioxide (hypothetical).

Tetroxides.—These contain four atoms of oxygen combined with the metal, thus:—

V_2O_4, vanadium tetroxide.
Sb_2O_4, antimony tetroxide.
Bi_2O_4, bismuth tetroxide.

Pentoxides.—These consist of molecules which contain five atoms of oxygen united by chemism with the electro-positive element, thus:—

Bi_2O_5, bismuth pentoxide.
Ta_2O_5, tantalum pentoxide.
N_2O_5, nitrogen pentoxide.
P_2O_5, phosphorus pentoxide.

Double Oxides or "saline oxides" may be regarded as the union of two oxides, thus:—

$$\underset{\text{Manganese monoxide.}}{MnO} + \underset{\text{Manganese dioxide.}}{MnO_2} = \underset{\text{Manganoso-manganic oxide, or Manganese sesquioxide.}}{Mn_2O_3}.$$

$$\underset{\text{Iron monoxide.}}{FeO} + \underset{\text{Iron sesquioxide.}}{Fe_2O_3} = \underset{\text{Ferroso-ferric oxide.}}{Fe_3O_4}.$$

In properties, oxides differ very much. Some readily give off oxygen when heated, lower oxides remaining. Others, gold and silver for example, part with their oxygen completely, the metal remaining. The oxides of some metals are volatile (ruthenium and osmium, examples) and can not be decomposed by heat. Others absorb oxygen when heated in the air, forming higher oxides.

Ic and Ous.—These terminations when applied to oxides show that the element forms at least two classes of oxides, *ous* terminating the lower and *ic* the higher, as:

MnO ..... Manganous oxide.
MnO_3 ..... Manganic oxide.
SnO ..... Stannous oxide.
SnO_2 ..... Stannic oxide.
CrO ..... Chromous oxide.
Cr_2O_3 ..... Chromic oxide.

A glance at the foregoing will show that oxides which terminate in *ous*, contain less oxygen in proportion to the metal than those which terminate in *ic*. This rule follows with the corresponding line of salts from each class of oxides.

Per, the prefix. This is applied to the oxide of an element which contains more oxygen than the basic oxide. The prefix *per* is old, having been once used by chemists

in contra-distinction to *proto*, where an element formed two classes of oxides that formed salts. *Proto*, in these cases, was applied to the oxide which contained the least oxygen, thus:—

Protoxide of iron was .... FeO.
Peroxide of iron was.....Fe_2O_3.

The prefix *proto* usually corresponded with our termination *ous*, meaning lower. The prefix *per* usually corresponded with our termination *ic*, and was applied to the higher class. It was customary to speak of the "sulphate of *proto*xide of iron" (now ferr*ous* sulphate), and "sulphate of *per*oxide of iron" (now ferr*ic* sulphate,) etc. (See *salts*.) For convenience the oxides have been divided into three classes.

Classes of Oxides.

These are *basic oxides*, *peroxides*, and *acid-forming oxides*.

Basic oxides, (*alkaline oxides*).—These neutralize acids and form compounds which are known as *salts*, thus:—

$$\underset{\text{Potassium monoxide.}}{K_2O} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Potassium sulphate.}}{K_2SO_4} + \underset{\text{Water.}}{H_2O}.$$

Again—

$$\underset{\text{Calcium oxide.}}{CaO} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Calcium sulphate.}}{CaSO_4} + \underset{\text{Water.}}{H_2O}.$$

Basic oxides also unite with water to form the class of bodies known as hydroxides or hydrates, thus:—

$$\underset{\text{Calcium monoxide.}}{CaO} + \underset{\text{Water.}}{H_2O} = \underset{\text{Calcium hydroxide.}}{Ca(OH)_2}.$$

$$\underset{\text{Potassium oxide.}}{K_2O} + \underset{\text{Water.}}{H_2O.} = \underset{\text{Potassium hydroxide.}}{2KOH}.$$

$$\underset{\text{Ferric oxide.}}{Fe_2O_3} + \underset{\text{Water.}}{3H_2O} = \underset{\text{Ferric hydroxide.}}{Fe_2(OH)_6}.$$

Hydroxides in turn are decomposed by acids to form salts with liberation of the water, thus:—

$$\underset{\text{Potassium hydroxide.}}{2KOH} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Potassium sulphate.}}{K_2SO_4} + \underset{\text{Water.}}{2H_2O.}$$

$$\underset{\text{Calcium monoxide.}}{Ca(OH)_2} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Calcium sulphate.}}{CaSO_4} + \underset{\text{Water.}}{2H_2O.}$$

It will be found that in many instances hydroxides are in practice used in preference to basic oxides, for making salts or the hydroxide only can be used as in making citrates of iron.

Peroxides (higher oxides or neutral oxides) contain more oxygen than basic oxides. They do not readily neutralize acids to give rise to salts, and their oxygen is in part given off when they are heated. As examples of *peroxides* we may name :—

BaO_2, peroxide (or dioxide) of barium.
MnO_2, peroxide (or dioxide) of manganese.
PbO_2, peroxide (or dioxide) of lead.

Acid-forming oxides (acid oxides) unite with water to form substances which are either acids or have acid properties, thus :—

$$\underset{\text{Sulphur trioxide.}}{SO_3} + \underset{\text{Water.}}{H_2O} = \underset{\text{Sulphuric acid.}}{H_2SO_4.}$$

$$\underset{\text{Nitrogen pentoxide.}}{N_2O_5} + \underset{\text{Water.}}{H_2O} = \underset{\text{Nitric acid.}}{2HNO_3.}$$

Examples of this class of oxides are shown by the acid-forming oxides of the nitrogen group of elements, and of the antimony group. We may state that this division of the oxides into three classes can not be accepted as a rule. There are oxides which may be either base or acid, and which under certain conditions act as the one, and again in the opposite manner. Thus while it is customary for chemists to apply the terms we have defined to indicate classes, it can not be considered that they are always appropriate. As examples of the exceptions we may cite manganese dioxide, MnO_2, or lead dioxide, PbO_2, each of which possesses both feebly basic and feebly acid properties, and both of which are classed with the neutral oxides. The student must bear in mind, therefore, that while we may, for convenience, refer to oxides as of these

three *divisions*, there is no dividing line. The oxides pass from the strongly alkaline potassium and sodium oxides to the strongly acid sulphur trioxide and phosphorus pentoxide, without a sharp line of demarkation. It is unfortunate that such a number of terms requires explanation. The steps through which our science has passed are marked by names that should be obsolete, but which the student of the present must consider.

Oxygen unites with carbon, hydrogen, sulphur, etc., and if the temperature be high enough such combination is accompanied with both light and heat. This is true of metals as well, and the phenomenon is termed *combustion.* Thus it is common to say that oxygen is a supporter of combustion, and that these bodies burn in oxygen. The fact is, however, that when the surfaces of the two bodies reach a certain temperature, combination ensues, and thus a jet of oxygen may be apparently made to burn in an atmosphere of hydrogen gas. The experiments with oxygen are usually given much prominence in school chemistries by reason of their brilliancy. Thus, a piece of fine iron wire tipped with sulphur, which is then ignited, burns with dazzling scintillations if it be thrust into a jar of oxygen. The result is black oxide of iron, Fe_3O_4.

If a small piece of phosphorus be lighted, and held in a jar of oxygen, combustion ensues apparently throughout the entire globe. This results from the combustion of the vaporized phosphorus. At last the globe is filled with dense white fumes of *phosphorus pentoxide* P_2O_5. This experiment is best performed in a bell glass of oxygen over water.

If a sliver of wood be lighted and the flame extinguished and the sliver with the red coal which remains be thrust into oxygen gas, flame is at once re-produced by reason of the heat generated from the increased chemical action. Under certain conditions oxygen assumes a form which is recognized by scientists as—

Ozone (*active oxygen*).—Formula, O_3.

This modification of oxygen has a peculiar odor and received the name *ozone* (I smell), from this fact. Ozone was discovered by Schönbein, in 1840, although previous to that (1785) Van Marum,† had noticed that oxygen gas through which a discharge of electricity had been passed would tarnish mercury and was possessed of an odor.

Ozone may be easily made by the slow oxidation of phosphorus, the experiment being best performed by placing in a glass jar water sufficient to partly immerse a stick of phosphorus, which is placed upon the bottom of the jar. Agitate the container occasionally, and in a short time the presence of ozone may be determined by the liberation of iodine from iodized starch paper.‡ Ozone is formed when a series of electrical sparks is passed through oxygen, and doubtless during thunder storms the electrical discharge is always productive of ozone in the atmosphere. When barium dioxide (BaO_2) is decomposed by means of sulphuric acid, the oxygen liberated contains a portion of ozone, thus:—

$$3H_2SO_4 + 3BaO_2 = 3BaSO_4 + O_3 + 3H_2O.$$

Sulphuric acid. Barium dioxide. Barium sulphate. Ozone. Water.

In reality but a small part of the oxygen thus freed is ozone, as the larger share of the reaction is expressed as follows:—

$$2H_2SO_4 + 2BaO_2 = 2BaSO_4 + O_2 + 2H_2O.$$

Sulphuric acid. Barium dioxide. Barium sulphate. Oxygen. Water.

Ozone is also produced and dissolved by many essential oils during the process of oxidation. Ozone has been accepted as a molecular form of oxygen containing three atoms (O_3) in each molecule. When ozone is formed the oxygen contracts by one-third its volume, and it is supposed that three molecules (or volumes) of oxygen unite to form two molecules (or volumes) of ozone according to the following equation:—

† R. and S., p. 194.

‡ This is properly iodide of potassium and starch paper, a process for the preparation of which is given under *tests* for ozone.

$$\underset{\text{Oxygen.}}{3O_2} = \underset{\text{Ozone.}}{2O_3}. \quad \text{Or,} \quad \underset{\text{Oxygen.}}{3O_2} = \underset{\text{Ozone.}}{2O_2O}.$$

This theory is farther supported by the fact that two volumes of ozone heated to 237° C. (458° F.) yield three volumes of oxygen. Other lines of investigation most carefully performed gave the same result, and we may accept as a fact that ozone is an allotropic form of oxygen, differing from ordinary oxygen in the number of atoms contained in each molecule. In consequence of this some choose to designate the *atom* of oxygen by the name oxygen, and to call the ordinary form of oxygen *oxide* of oxygen, OO, and to thus derive the name *peroxide* (higher oxides) of oxygen for ozone, OO_2.

Properties.—Ozone has a strong and very peculiar odor, is one of the most powerful of oxidizers, and will even destroy gutta-percha. It parts with one atom of oxygen during oxidation, the *bulk* of the gas remaining unchanged.† Blood is decolorized by ozone, albumen is destroyed, and most metals are oxidized, while alcohol C_2H_6O, is changed to aldehyde C_2H_4O. It is slightly soluble in water, and quite soluble in certain volatile oils. When hydrogen dioxide (H_2O_2) is mixed with ozone, both bodies decompose at once, oxygen and water resulting, thus:—

$$\underset{\text{Ozone.}}{O_3} + \underset{\text{Hydrogen dioxide.}}{H_2O_2} = \underset{\text{Water.}}{H_2O} + \underset{\text{Oxygen.}}{2O_2}.$$

This fact, until explained, led certain chemists to believe in a form of oxygen which could unite with ozone to form ordinary oxygen. To this the name *anti-ozone* was applied, but the term is now obsolete as the so-called antozone was shown to be hydrogen dioxide.

Test.—Ozone liberates iodine from iodide of potassium, consequently it will change to blue, paper which is moistened with solution of iodide of potassium in starch water. To prepare this paper, mix one part of starch with 40 parts of water and bring the mixture to a boil. Then dissolve in it one part of potassium iodide, and wet sheets

† This corresponds to the law of combination of gases by volume, and it will be observed that the production of ozone from ordinary oxygen also agrees with that law.

of white porous paper with the solution and dry. When this paper is moistened, ozone will change the color of the paper to blue.

Oxyhydrogen Flame.—When oxygen and hydrogen unite water is produced, (see water). The combustion of pure hydrogen in the air gives rise to this compound only. The act of combination results in the liberation of much heat with but little light. When a jet of oxygen gas is thrown into a jet of burning hydrogen, the most intense heat results, and this produces the *oxyhydrogen flame.* In practice the two are conducted to a platinum tip or nozzle and they mix together as they issue from the orifice. The hydrogen is first turned on and ignited, then the oxygen is admitted, thus overcoming danger of explosion. The apparatus is known as the *oxy-hydrogen blowpipe,* and within the pale blue flame iron burns like tinder and even platinum melts as though it were wax. If a piece of lime (calcium monoxide) be placed in the flame, it assumes an intensely incandescent state, and the light is known as the *oxyhydrogen* or Drummond light. Coal gas is often used in place of hydrogen, but this produces an inferior light. Sometimes the jet of oxygen is thrown through the flame of an alcohol lamp; this in common language is distinguished as the *oxycalcium light.*

SULPHUR.

Symbol, S. Atomic weight, 31.98.

Sulphur is found native in volcanic neighborhoods. The compounds of sulphur with iron (FeS_2) and with copper and iron ($CuFeS_2$) are known as iron and copper pyrites. Galena (PbS) is a compound of sulphur and lead; cinnabar (HgS) is a compound of mercury and sulphur; and blend, (ZnS) of zinc and sulphur. The compounds of sulphur and metals, and sulphur, metals and oxygen, occur over the entire globe, often in great quantities. Sulphur is mostly obtained from Italy and Sicily, where it is found in impure masses throughout the volcanic districts. To extract it, excavations in the ground are filled with the ore and a fire is applied to the top of each

heap. This fuses a part of the sulphur which collects at the bottom of the hole and is ladled out, the result being crude sulphur.

Refined Sulphur. Brimstone.

This is made by fusing crude sulphur and pouring it into moulds. This in commerce is also known as *roll sulphur*. If the liquid sulphur in mass is permitted to cool, the broken lumps are known as *block sulphur*.

Flowers of Sulphur.

This is the ordinary sulphur of trade. It is made by heating crude sulphur to the boiling point, when it vaporizes and the vapor is led into a cool chamber where condensation ensues, a light yellow crystalline powder resulting. The name *flowers* of sulphur was derived from the regular and beautiful form of the figures produced by the condensed vapor, after the manner of the frost figures (flowers) upon windows. A common error is that of calling this form of sulphur, *flour* of sulphur. Sulphur is dimorphous, forming crystals of two distinct systems, the rhombic and the monoclinic, and it assumes many crystalline modifications of these. Sulphur is lemon yellow, tasteless and odorless. It melts at 114.5° C. (Brodie), forming a yellow solution and this solution solidifies at the same temperature; or it may remain fluid until cooled to 111.5° C. Sulphur boils at 440° C, and the vapor has a red color. It dissolves freely in carbon disulphide, in certain oils, and in fats, and slightly in alcohol, ether, and in benzine. If sulphur be melted and the heat be gradually increased the liquid changes in color to red, and when the temperature reaches 200° C., or 250° C, it becomes almost black, and thickens. If this be now poured into cold water the result is a plastic, elastic mass known as *amorphous* or soft sulphur. This is not permanent, but in a few days becomes opaque, yellow and brittle. Soft sulphur is known as an *allotropic* form of sulphur. Sulphur is used in the arts and manufactories in immense amounts. It is present in combination, as sulphate of calcium, potassium, etc., in all soils, and its

compounds are found in both the animal and the vegetable kingdom. Sulphur unites with hydrogen to form *sulphide of hydrogen* (sulphuretted hydrogen) or mono-sulphide of hydrogen ; H_2S. This is described by us under *acids.* (See hydrosulphuric acid, page 102)

Hydrogen Persulphide, H_2S_2.

This is the only other known compound of hydrogen and sulphur, it is a yellowish oily body of no use in medicine.

Sulphur and Chlorine.

There are three compounds of sulphur and chlorine, but neither is of value in a medicinal sense. The monochloride is used as a solvent of sulphur in certain branches of the arts. These compounds are :

Sulphur monochloride, S_2Cl_2.
Sulphur dichloride, SCl_2.
Sulphur tetrachloride, SCl_4.

One compound of sulphur and bromine is known (S_2Br_2) but it is of no practical use.

Sulphur Moniodide. Iodide of Sulphur, S_2I_2.

This is a gray solid of a metallic appearance, and which breaks with a crystalline fracture. It is easily made by warming in a flask a mixture of four parts of iodine and one part of sulphur (U. S. P.) Iodide of sulphur evolves iodine when exposed to the air. If it be boiled with water, complete decomposition ensues and sulphur remains. Iodide of sulphur must be kept closely corked, in small vials.

Sulphur and Oxygen.

Sulphur unites with oxygen to form three oxides :

Sulphur dioxide (sulphurous anhydride or sulphurous acid gas), SO_2.
Sulphur trioxide (sulphuric anhydride), SO_3.
Sulphur sesquioxide, S_2O_3.

Two of these compounds are of great interest, inasmuch

as each by combining with a molecule of water, will give rise to an important acid, thus:

Sulphur dioxide, $SO_2 + H_2O$ = Sulphurous acid, H_2SO_3.
Sulphur trioxide, $SO_3 + H_2O$ = Sulphuric acid, H_2SO_4.

The formation and chemistry of both of these acids have been discussed under *acids* of this work. *Sulphur* sesquioxide is of no practical use, and it is not necessary for us to describe it.

Tests for Sulphur.—Mix the compound with a little pure carbonate of sodium, and fuse on a piece of charcoal or platinum, by means of the blowpipe flame. Dissolve the mass in a little water, filter it, and warm in contact with a bright silver coin. If sulphur be present in the original compound, the coin will turn brown or black.

Soluble Sulphides.—If hydrochloric acid be added to the solution of a sulphide which can be decomposed by that reagent, *sulphide* of *hydrogen* will be evolved, and this will blacken a strip of paper moistened with solution of lead acetate and suspended over the vessel. If the sulphide is not acted on by hydrochloric acid, add a piece of *pure zinc*, the sulphide of hydrogen will be at once evolved.

Insoluble Sulphides are to be finely powdered, and fused with a mixture of nitrate of potassium and carbonate of sodium. This dissolved in water, filtered, and the filtrate acidulated with hydrochloric acid, yields a solution of *sulphuric acid.* If a solution of barium chloride be now added, sulphate of barium precipitates. (See sulphuric acid.)

SELENIUM.

Symbol, Se. Atomic weight, 78.0.

Berzelius discovered this element in 1817. Selenium is found in small amounts, rather widely distributed in certain minerals, as iron pyrites, copper pyrites and as selenide of sulphur. It is obtained from the red deposit found in the chambers of apparatus where iron and copper pyrites are used in making oil of vitriol. Selenium assumes two forms. One is insoluble in carbon disul-

phide, and is known as *metallic* selenium. The other dissolves in carbon disulphide, has no definite melting point, and is known as amorphous selenium. Metallic selenium melts at 217° C, and when cooled assumes the amorphous form. It conducts electricity, and when it is heated in the air burns with a blue flame, and evolves a disagreeable odor resembling "rotton horse radish." None of the compounds of this element are used in medicine.

TELLURIUM.

Symbol, Te. Atomic weight, 128.00.

Tellurium is one of the rare elements. Richenstein, in 1782, examined the native ore and suggested the probability that it contained a new metal. His conclusion was afterward supported by Klaproth (1798). This element occurs native, and in combination with metals. The process employed to separate it is of little interest to pharmacists. Tellurium crystallizes in form of rhombohedra, has a metallic lustre, a bluish-white color, and is so brittle as to readily powder in a mortar. It burns with a blue flame, dissolves in cold fuming sulphuric acid, and it rapidly oxidizes in nitric acid. Tellurium has the specific gravity 6.24 and melts at 500° C. It may be distilled in an atmosphere of hydrogen. The compounds of tellurium with other elements are of no use in medicine, and do not require notice from us.

NITROGEN.

Symbol, N. Atomic weight, 14.01.

Nitrogen was named by Chaptal, in consequence of the fact that it is a component of saltpetre (nitrum). It was identified previous to that, by Rutherford, (1772), as a constituent of the air. Nitrogen is easily prepared by inverting a bell glass over a basin of water on the surface of which a small vessel containing a piece of ignited phosphorus is floating. The phosphorus burns and forms phosphorus pentoxide, (P_2O_5), which exhibits itself as dense white fumes. These are afterward absorbed by the water. When the gas within the glass has become color-

less, it is almost pure nitrogen, and the gas will occupy four-fifths the space originally filled by the air. Nitrogen may be obtained otherwise, but the above is most simple.

Nitrogen is a colorless, odorless and tasteless gas. It does not support combustion and combines with but few elements in a direct manner. Its compounds formed indirectly, are numerous. Some are the most violent of explosives, others valuable in both the arts and in medicine. Nitrogen forms four-fifths the bulk of the atmosphere, serving as a diluent of the oxygen, the remaining fifth. Its specific gravity is 0.9713. Water dissolves it less readily than it does oxygen, and therefore the proportion of oxygen in water is greater than that of oxygen in the air. The albuminous bodies of both the animal and the vegetable kingdoms contain nitrogen as an essential constituent.

Nitrogen and hydrogen unite to form ammonia (NH_3), but we shall study this body with the potassium group of elements to which the reader is referred.

Nitrogen and chlorine unite to form a liquid which is a most dangerous explosive, *Dulong*, its discoverer, losing an eye and three fingers while experimenting upon it. A small glass tube containing a few grains of the liquid also exploded in Faraday's hand, and the shock rendered him unconscious. Sir Humphrey Davy also experienced a fearful explosion and received some injury. Chloride of nitrogen consequently has not been analyzed and it is not used in medicine. *Bromide* of *nitrogen* is also an explosive, and but little known. The *iodides* of nitrogen are better known, but they explode violently even when touched.

Nitrogen and Oxygen.

There are five oxides of nitrogen; as follows:—

Nitrogen monoxide (nitrous oxide) ..... N_2O.
Nitrogen dioxide (nitric oxide) ..... N_2O_2 or NO.
Nitrogen trioxide ..... N_2O_3.
Nitrogen tetroxide (nitrogen peroxide) ..... N_2O_4 or NO_2.
Nitrogen pentoxide ..... N_2O_5.

None of the above are used in medicine. Three unite, severally, with a molecule of water to form acids as follows:—

$$\underset{\text{Nitrous oxide.}}{N_2O} + \underset{\text{Water.}}{H_2O} = \underset{\text{Hyponitrous acid.}}{2HNO}.$$

$$\underset{\text{Nitrogen trioxide.}}{N_2O_3} + \underset{\text{Water.}}{H_2O} = \underset{\text{Nitrous acid.}}{2HNO_2}.$$

$$\underset{\text{Nitrogen pentoxide.}}{N_2O_5} + \underset{\text{Water.}}{H_2O} = \underset{\text{Nitric acid.}}{2HNO_3}.$$

The foregoing acids are of interest and value as they enter into combination with, or result from the decomposition of many substances which are used in medicine. They have been studied under the section of our work devoted to acids.

The compounds formed by the union of nitrogen with single elements are known as *nitrides*.

PHOSPHORUS.

Symbol, P. Atomic weight, 30.96.

It is accepted that phosphorus was discovered by Brant, in the year 1669. Phosphorus is very widely distributed, always in combination. It is found in the ocean, the waters of springs and rivers, and in all soils. It is an essential constituent of the entire vegetable kingdom, and phosphate of calcium is the larger part of the bones of animals. It is likewise an indispensible component of muscle and of nerve tissue, and is intimately connected with the vital functions. It is unnecessary for us to give the process employed in making phosphorus, as this element is never prepared on a small scale. It is produced from burnt bones, but the first phosphorus was derived by its discoverer from urine.

Phosphorus is a remarkable element. It assumes several allotropic forms. The ordinary phosphorus is waxy at common temperatures and is of a translucent, yellowish color. It reaches market in cylindrical pieces, half an inch or so in diameter and of from four to six inches in length. It melts at 44.3° C. (Schrotter), and will not

conduct electricity. Phosphorus is insoluble in water, slightly soluble in ether, turpentine and essential oils. Carbon disulphide dissolves from seventeen to eighteen times its weight of phosphorus.

From solution in this substance, by slow evaporation, phosphorus separates in crystalline form. If the solution of phosphorus in carbon disulphide be thrown upon a surface which is exposed to the air, the carbon disulphide evaporates, and the phosphorus inflames spontaneously, therefore, great care must be taken that such a mixture is well preserved. Phosphorus is very inflammable, igniting simply upon exposure to the atmosphere. It must always be preserved beneath water, and when desired in small pieces it should be invariably cut beneath the surface of water, and it should *never* be handled in the air as the heat of the hand is sufficient to inflame it. Phosphorus produces an intensely hot flame, and this is very difficult to extinguish. When phosphorus is exposed to the air in the dark it appears luminous. This is true even if a *very* small portion of oxygen is present. In *pure* oxygen it remains unchanged for some time, and is said not to undergo the least oxidation. The foregoing fact has not been explained. Common phosphorus preserved under water becomes covered with a white crust; this is ordinary phosphorus corroded by oxidation until it assumes the opaque white color. Phosphorus is now manufactured in immense amounts and is consumed in various ways, chiefly for tipping lucifer matches. It is a violent poison, and there is no chemical antidote. The treatment consists in the administration of oils, emulsions and mucilaginous drinks.

Amorphous or Red Phosphorus.

This is a form of phosphorus produced by heating the ordinary article in a close vessel to the temperature of 300° C. The change may also be induced by other methods,

† The term amorphous is applied to bodies that are destitute of regular structure. Several elements, as phosphorus, carbon, and selenium, assume a crystalline and an amorphous form.

but the foregoing is the most practical. Amorphous phosphorus has a dark reddish-brown color and possesses a metallic lustre. It is tasteless and odorless when pure, but as found in market has usually a distinct odor due to impurities. This form of phosphorus (pure) does not oxidize upon exposure to the air, will not ignite by friction, and is said to be perfectly harmless even if swallowed in large amount. It is insoluble in liquids which freely dissolve the ordinary article. The phosphorus of commerce is impure from the presence of arsenic. Phosphorus also assumes a form called *metallic* or rhombohedral phosphorus. This is made by heating phosphorus and lead together in sealed tubes, and has a lustrous crystalline appearance.

Hydrogen and Phosphorus unite in three proportions:

PH_3, hydrogen phosphide, phosphoretted hydrogen, phosphine,
P_2H_4, liquid hydrogen phosphide.
P_4H_2, solid hydrogen phosphide.

The preparation of phosphide of hydrogen (PH_3) is an old and common lecture-room experiment. It is of no particular value to the student of medicine, therefore, we refer such as are interested in experiments for the sake of entertainment only, to works devoted to the study of general chemistry.

Phosphorus forms with chlorine—

PCl_3, phosphorus trichloride.
PCl_5, phosphorus pentachloride.

With bromine phosphorus forms:—

PBr_3, phosphorus tribromide.
PBr_5, phosphorus pentabromide.

With iodine phosphorus forms:—

P_2I_4, phosphorus di-iodide.
PI_3, phosphorus tri-iodide.

There is one compound of phosphorus and fluorine:

PF_5, phosphorus pentafluoride.

Phosphorus and oxygen unite in two proportions:

Phosphorus trioxide (phosphorus anhydrid), P_2O_3.
Phosphorus pentoxide (phosphoric anhydrid), P_2O_5.

These unite with water to form interesting acids.

Phosphorus Trioxide, P_2O_3.

This is made by causing phosphorus to undergo slow combustion, a feat accomplished by placing a piece of phosphorus in a glass tube, heating it and then passing over it a slow current of air. The phosphorus burns slowly and *phosphorus trioxide* results. This is deposited in the cool part of the tube as a white powder. It has a garlic-like odor and unites with water to form phosphorous acid, thus:—

$$\underset{\text{Phosphorus trioxide.}}{P_2O_3} + \underset{\text{Water.}}{3H_2O} = \underset{\text{Phosphorous acid.}}{2H_3PO_3}.$$

Phosphorous acid is not used in medicine and is considered poisonous. It gives rise to the *phosphites*.

Phosphorus Pentoxide, P_2O_5.

This is the result of burning phosphorus in air or oxygen. In our experiment upon nitrogen (page 147) this compound is formed. It may be made by inverting a bell glass over a piece of phosphorus which is burning on a plate. White fumes result, and a white flocculent powder condenses upon the cool parts of the plate and upon the bell glass. Occasionally a small portion of trioxide forms during the latter part of the process, when the air has become mostly depleted of oxygen. This is distinguished by the garlic-like odor before mentioned. Phosphorus pentoxide is a white, odorless powder, and will not affect litmus. It has an *intense* affinity for water, abstracting that substance from most other substances which contain it; even abstracting the elements of water (oxygen and hydrogen) from bodies that contain them. This combination with water forms *meta-phosphoric acid* (see acids), as follows:—

$$\underset{\text{Phosphorus pentoxide.}}{P_2O_5} + \underset{\text{Water.}}{H_2O} = \underset{\text{Meta-phosphoric acid.}}{2HPO_3}.$$

The medicinal substances which contain phosphorus in combination are treated of under each element. The compounds of phosphorus and sulphur, and of phosphorus and nitrogen are of interest to the student of general chemistry, but can not be considered in this work.

ARSENIC.

Symbol, As. Atomic weight, 74.9.

Arsenic is generally obtained from arsenical iron pyrites (mispickel, Fe S As). It exists in many minerals, and is found native. Orpiment and realgar are *sulphides* of arsenic. Arsenic resembles in appearance the metals antimony or bismuth, but has properties which class it with the non-metallic elements. It has a steel-gray color and a metallic lustre. Its specific gravity is 4.710, and when it is heated in the air, it at once assumes the form of gas. If it be heated (Landolt) under pressure, however, it can be melted. Its vapor has the odor of garlic. Arsenic unites with non-metallic bodies and with most metals. It oxidizes in moist air, and burns in oxygen with a white flame, arsenic trioxide resulting.

Arsenic is very poisonous, and yields compounds, most of which are fearfully destructive to life. The most important and common compound of arsenic is the white arsenic of commerce (arsenic trioxide).

Arsenic and Hydrogen unite to form two compounds:

AsH_3, hydrogen arsenide, arsine, or arseniuretted hydrogen.

As_2H_2, solid hydrogen arsenide.

The latter substance is of no use, but arseniuretted hydrogen is very important, as it constitutes one of the most delicate tests for arsenic.

Arseniuretted Hydrogen, AsH_3.

This gas is always formed when hydrogen gas is liberated from any liquid which contains arsenic in solution. It is made pure by decomposing arsenide of zinc, As_2Zn_3, with dilute sulphuric acid. If a few grains of arsenious

anhydrid be mixed with water, to which some pieces of metallic zinc and a little dilute sulphuric acid are added, a mixture of hydrogen gas and arseniuretted hydrogen result. Arseniuretted hydrogen is very poisonous, therefore care must be taken that the gas be not inhaled. When it is passed into the solution of a gold or silver salt, the metal is precipitated. Arseniuretted hydrogen has a garlic-like disagreeable odor. If it be conducted through a glass tube heated at a given point with an alcohol flame, the gas is decomposed and a precipitate of a metallic appearance is produced in the cool part of the tube.

If a stream of arseniuretted hydrogen be ignited as it issues from a glass tube, and a cold porcelain plate be held in the flame, a metallic, mirror-like spot will result.

The foregoing is "*Marsh's Test for Arsenic,*" and it is so delicate that $\frac{1}{70000}$ part of a grain (Otto) of arsenic can be identified.

Marsh's Test for Arsenic.†

This is to be conducted as follows: Into the four ounce flask (fig. 50) A, pour two ounces of water, and place in it a few scraps of pure metallic zinc. Then add the substance which is suspected to contain arsenic and

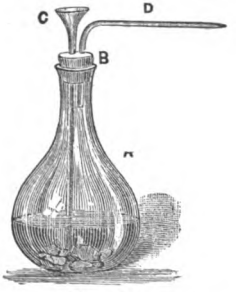

FIG. 50.

† Beginners in chemical work should not attempt to analyze the contents of a stomach which is suspected to contain arsenic. Few chemists, even with great experience in synthetical work, care to assume the responsibility of such examinations. It is well, however, that the student should experiment upon the detection of arsenic in

insert the cork B. Through this cork a funnel tube, C, extends, the lower end of which enters the water. The glass tube D, extends through the cork and this is bent as represented by the figure and drawn at its extremity to a point.

Now pour two fluid drachms of sulphuric acid through the funnel tube C, and when effervescence begins heat the tube D, by means of a spirit lamp as represented in our cut (fig 51), and cool the extremity of the tube by means of a

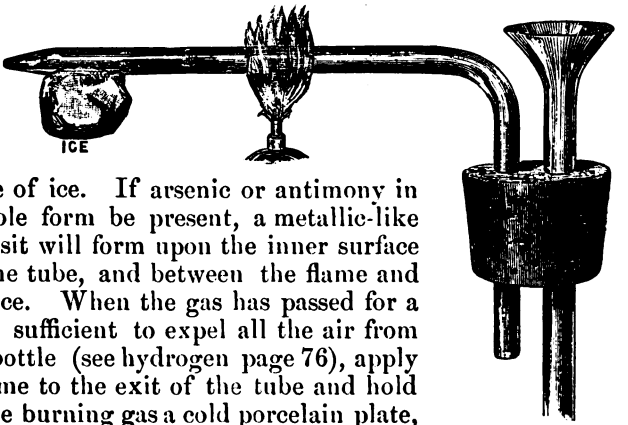

FIG. 51.

piece of ice. If arsenic or antimony in soluble form be present, a metallic-like deposit will form upon the inner surface of the tube, and between the flame and the ice. When the gas has passed for a time sufficient to expel all the air from the bottle (see hydrogen page 76), apply a flame to the exit of the tube and hold in the burning gas a cold porcelain plate, or the top of a white porcelain jar (fig. 52). The formation of a metallic-like stain corroborates the foregoing indication of arsenic or of antimony. To determine the nature of the metallic mirror, pass sulphide of hydrogen through the tube D, gently heating the deposit with a spirit lamp. It will assume a yellow (arsenic) or an orange (antimony) color, then will volatilize and recondense as a yellow (arsenic) or an orange (antimony) colored sublimate. This is either sulphide of arsenic, or sulphide of antimony. Now adapt the cork, containing the tube, to a flask in

such medicinal substances as are liable to be contaminated with that body. Cases of poisoning, however, where the future of a human being may depend upon the report of the chemist, should be investigated by only such as are prepared with reagents of undoubted purity, and such as have great experience in analytical work.

which some common salt has been placed, and into which just previous to the insertion of the cork a little sulphuric acid was poured. Hydrochloric acid gas will be generated, and if the precipitate within the tube be sulphide of arsenic no change will result. If upon the other hand it be sulphide of antimony, solution results and the orange color disappears. There are other tests for arsenic, but

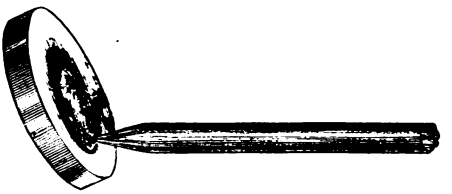

FIG. 52.

it is unnecessary for us to consume time with them. Remember, that both metallic zinc and sulphuric acid are liable to be contaminated with arsenic, so that the detection of arsenic by Marsh's test implies simply that arsenic is present in some of the materials. In investigations where reputation or even life depends upon the result, reagents of *undoubted* purity must be employed.

Arsenic and Chlorine.

These unite to form arsenic trichloride, $AsCl_3$, a very poisonous liquid. With bromine arsenic forms a tribromide, $AsBr_3$. Neither of the above is employed in medicine.

Arsenic Tri-iodide, or *Iodide of Arsenic*, AsI_3.

This is a red crystalline body when pure, but as made by the U. S. P. consists of an orange-red brittle substance which breaks with a crystalline fracture. It is made (U. S. P.) by mixing in a mortar sixty parts of powdered arsenic (metallic) with three hundred parts of iodine; the mixture is then to be placed in a test-tube or a flask, and fused. Care must be taken that the heat be not too great, and that the operator does not inhale the fumes which escape during the operation.

Arsenic and Oxygen unite to form—

Arsenic trioxide, As_2O_3.
Arsenic pentoxide, As_2O_5.

The first of these bodies is the common arsenic of commerce, or *arsenious anhydrid.* It is often called *arsenious acid,* but in reality it is an acid-forming oxide. When the trioxide is dissolved in water, it is accepted that arsenious acid results; thus:—

$$\underset{\text{Arsenic trioxide.}}{As_2O_3} + \underset{\text{Water.}}{3H_2O} = \underset{\text{Arsenious acid.}}{2H_3AsO_3}.$$

This acid has never been prepared pure, (see arsenious acid, page 87.)

Arsenic pentoxide, As_2O_5, or arsenic anhydrid, is the acid-forming oxide which gives rise to arsenic acid. It is best made by heating arsenic acid to dull redness, whereby water is driven off as follows:—

$$\underset{\text{Arsenic acid.}}{2H_3AsO_3} = \underset{\text{Arsenic pentoxide.}}{As_2O_5} + \underset{\text{Water.}}{3H_2O}.$$

A very high heat decomposes it into arsenic trioxide and oxygen, thus:—

$$\underset{\text{Arsenic pentoxide.}}{As_2O_5} = \underset{\text{Arsenic trioxide.}}{As_2O_3} + \underset{\text{Oxygen.}}{O_2}.$$

For descriptions of the acids of arsenic, see *acids.*

Arsenic and Sulphur.

There are three sulphides of arsenic, viz:

As_2S_2, arsenic disulphide, or realgar.
As_2S_3, arsenic trisulphide, or orpiment.
As_2S_5, arsenic pentasulphide.

The first and second of these bodies are found native, and have been used since an early day in many branches of the arts, orpiment especially having been employed as a pigment under the name *King's yellow.* It is very poisonous, however, and has been replaced by chrome colors.

The trisulphide (As_2S_3), and the pentasulphide (As_2S_5), of arsenic form an interesting series of acids and salts, known as thio-arsen*ites* and thio-arsen*ates.* They are of

no use in medicine. however. Arsenic forms with phosphorus the compound arsenic phosphide, AsP. Arsenic and its compounds are very poisonous. The antidotes, and the treatment of such unfortunates as may chance to require aid, have been discussed under acids, (page 88). The compound of Arsenic and a metal is an *arsenide.*

BORON.

Symbol, B. Atomic weight, 11.00.

Boron is found in natural combination, as boric acid or as a salt of this acid. It is remarkable as being the only non-metallic element which has refused to combine with hydrogen. Boron exists in two allotropic modifications, *amorphous boron,* and *crystallized* or *adamantine* boron. Neither possesses any particular interest to the student of medicine.

Boron and Chlorine.

Boron unites with chlorine, bromine and with fluorine to form boron trichloride, BCl_3, boron tribromide, BBr_3, and boron trifluoride, BF_3. Neither is employed in medicine.

Boron and Oxygen.

Boron trioxide, B_2O_3, the only known oxide of this element, results from the combustion of boron in oxygen gas, or from heating boric acid to redness. It unites with water to form boric acid, $B(OH)_3$, thus:—

$$\underset{\text{Boron trioxide.}}{B_2O_3} + \underset{\text{Water.}}{3H_2O} = \underset{\text{Boric acid.}}{2H_3BO_3}.$$

Boric acid or salts of boric acid, are the compounds of this element mostly used in the arts and in medicine. Boric acid has been described (page 90), and the compounds of boron with the metals will be considered with each element.

Boron unites with nitrogen to form *boron nitride* BN, and with sulphur to form boron trisulphide.

Tests for Borates.—Add hydrochloric acid to the solution of an alkaline borate until the acid is in slight excess,

then dip a piece of tumeric paper into it and dry at the temperature of 100° C. (212° F.) The paper will now have a peculiar reddish tint best shown by placing it beside a piece of unchanged tumeric paper.

2d. Add a small portion of a powdered borate to strong sulphuric acid, and then mix with it a little alcohol. If the alcohol be now ignited the edges of the flame will appear greenish. Chlorides interfere with this test and copper salts impart a like green color.

SILICON (*Silicium*).

Symbol, Si. Atomic weight, 28.0.

This element is very abundant in combination, forming a large part of the earth's crust, and a large share of portions of certain plants, as the bark of hickory and the glassy surface of straw. It is always united with oxygen, SiO_2. Amorphous silicon is a dark brown powder. A crystalline modification may be prepared. The amorphous form will dissolve in solution of hydrofluoric acid; other acids have no effect upon it. When it is heated in the air oxygen combines with it, dioxide of silicon resulting. The crystallized modification is dissolved by a mixture of nitric and hydrofluoric acids. It slowly oxidizes if heated to redness in oxygen.

Silicon and Hydrogen unite to form silicon hydride, SiH_4.

Silicon and Chlorine form silicon tetrachloride, $SiCl_4$, and silicon trichloride, Si_2Cl_6. Two compounds of silicon with bromine, and two with iodine are known, neither being used in medicine. They correspond with the chlorine compounds.

Silicon and Oxygen form silicon dioxide (SiO_2) or silica. This is an abundant mineral, and immense deposits almost pure are found in different parts of the world. Quartz and white sand are familiar examples of native silica. There are three native forms of silica, and many minerals contain one or more of these modifications. The abundance of silica may be understood, when we consider

insoluble in all the chemical reagents,† can not be volatilized, and has never been fused. It is a rapid absorber of gases and vapors, and will abstract many uncrystallizable substances from solution and hold them. This principle is employed in obtaining certain bitter principles from plants, and in purifying water. If a piece of cane sugar, or a bit of wood, be covered with sulphuric acid, disintegration results, the acid abstracts the elements of water (oxygen and hydrogen) from the organic body, and a black mass (impure carbon) remains.

Lampblack is a form of finely divided carbon. It is made when substances rich in carbon, such as wax, fat, rosin, etc., are burned in a limited supply of air. The heat of the flame decomposes the inflammable material, but the oxygen is not present in sufficient amount to consume the carbon. The consequence is that the carbon of the *smoke* condenses and the deposit is finely divided carbon. Lampblack is also formed by the incomplete combustion of coal gas, and we have known of its practical production from natural gas of the oil regions in this country. Large amounts of lampblack are produced by burning corn cobs previously saturated with rosin.

Animal Charcoal.—This form of carbon contains considerable nitrogen and all of the inorganic matter of bones. It is made by heating bones in iron retorts to redness. The distillate in this case is known as animal oil, or *oleum animale Dippelii* (Dippel's oil). Animal charcoal is of great value in sugar refining. It is known as bone-black.

Purified animal charcoal is made by digesting ordinary animal charcoal in hydrochloric acid. This removes the carbonate and phosphate of calcium, and thus fits the charcoal for decolorizing acid liquids. All forms of carbon are odorless and tasteless, and none of the modifications can be dissolved in any known liquid at ordinary tempera-

† Melted iron combines with carbon to form a *carbide* of iron. It is probable that melted iron *dissolves* carbon as it will take up four per cent. of that element. When the iron cools, a portion of the carbon separates as graphite.

tures. Carbon is not affected by air or water. Grains of corn which were charred in the expedition made upon the Indians of Western New York remain perfect to the present day. Organic matter, both animal and vegetable, contains a large amount of carbon in combination. Plants derive this carbon from air, and from the water of the soil, in which it exists as carbon dioxide or carbonic acid. The gas is decomposed in the plant tissue, carbon is fixed and oxygen is liberated as a gas. Animals feed upon plants and thus transfer the carbon to their systems. The study of the carbon compounds is intimately connected with organic bodies. In organic chemistry (so-called) we expect to consider the most important of the carbon compounds, for all of the alkaloids, the vegetable acids, the glucosides, etc., etc., contain this element. Indeed, it is frequently stated that organic chemistry is "The study of carbon compounds and of the migration of the carbon atoms," and the field is a large one. The study of carbon in its connection with organic bodies is the great work of the foremost chemists of the age. It must not be expected that a book of this description will more than touch the subject. We can only notice, here and there, a compound that chances to be connected with the study of the pharmacist.

Carbon combines with hydrogen to form a series of compounds called *hydrocarbons*. These are all volatile, and generally result from the decomposition of organic bodies. Their construction is complex and the number is undetermined. A few of them might with propriety be studied here, but we shall defer notice of such for part second of this work.

Carbon and Chlorine.

Carbon and chlorine will not unite directly. Combinations of carbon and chlorine may be made, however, by decomposition of certain hydrocarbons whereby chlorine is substituted for hydrogen.

Carbon and Oxygen.

Carbon combines with oxygen to form two oxides:

Carbon monoxide, CO. Carbon dioxide, CO_2.

These are both gases. Carbon monoxide or carbonic oxide results when zinc oxide and charcoal are heated together. It may be made by other processes, but as this gas is not used in medicine or the allied sciences it is unnecessary to consume time with them. Carbon monoxide is colorless, tasteless, and almost insoluble in water. It burns with a blue flame, carbon dioxide resulting, thus:—

$$\underset{\text{Carbon monoxide}}{2CO} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Carbon dioxide.}}{2CO_2}.$$

Carbon monoxide combines with the hæmaglobin of the blood, and is very destructive to animal life if inhaled.

Carbon dioxide (CO_2) is the familiar carbonic acid or carbonic acid gas of the past. It results from the combustion of carbon in the air and is in reality the acid-forming oxide which gives rise to carbonic acid (see acids), thus:

$$\underset{\text{Carbon dioxide.}}{CO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Carbonic acid.}}{H_2CO_3}.$$

Carbon dioxide exists free in the air to a slight extent. It is easily prepared by pouring hydrochloric acid upon pieces of marble. Carbon dioxide is the gas which results when an acid is added to a carbonate. Carbon dioxide is formed during fermentation, and often accumulates in wine and beer cellars to such an extent as to displace the air. This is the gas which accumulates, in unused wells, and is termed in common language *choke-damp* or *damps*.

Carbon dioxide is colorless, and produces a feeling of strangulation if inhaled even in small amounts. It will neither support combustion nor respiration, flame is at once extinguished, and animals quickly die in it. It is heavier than the air, having the specific gravity 1.5241 (Regnault), and it may be poured from one vessel to another. This gas may be liquefied by both pressure and a low temperature, forming a colorless liquid. Carbon dioxide is very soluble in water and the so-called soda-water, used as a beverage, is a solution of this gas. The sharp taste of fermenting wine and cider is derived from the presence of carbon dioxide.

Carbon and Sulphur unite to form *carbon disulphide*, CS_2, or bisulphide of carbon. Sometimes this preparation is called by the old name, bisulphuret of carbon. This is made on a large scale by passing the vapor of sulphur over red-hot charcoal, the impure product being condensed and sold for uses in the arts.

Carbon disulphide of commerce has an offensive odor, owing to the sulphur compounds which are associated with it. When pure it is a colorless, very refractive liquid, with a pleasant ethereal, somewhat anise-like odor. It is inflammable, and its vapor mixed with atmospheric air explodes with violence if lighted. It is poisonous and the inhalation of its vapor produces serious effects. Carbon disulphide is a solvent of oils and gutta percha, and is used in the arts for this purpose.

Carbon and Nitrogen unite to form cyanogen gas C_2N_2 or Cy_2, in reality dicyanogen, a colorless gas. It may be made by heating cyanide of mercury to dull redness in a glass tube, metallic mercury and cyanogen being liberated, the decomposition resulting as follows:—

$$\underset{\text{Mercuric cyanide.}}{2Hg(CN)_2} = \underset{\text{Mercury.}}{2Hg} + \underset{\text{Cyanogen gas.}}{2C_2N_2}.$$

Cyanogen gas has an odor which resembles peach pits. It is very poisonous, has the specific gravity of 0.866, and may be liquefied by both pressure and a low temperature.

Cyanogen (Monad radical), CN or Cy, is the acidulous radical of hydrocyanic acid and of all the cyanides. It is capable of replacing one atom of hydrogen, or of combining with one atom of hydrogen. It may be compared with chlorine in this respect, as the cyanides very much resemble the chlorides. The radical of cyanogen thus plays the part of a single atom; and when cyanogen is in a free state two of these radicals are united, thus:—

$$CN + CN = C_2N_2. \quad \text{Or } Cy + Cy = Cy_2.$$

Consequently, we consider the *real* molecule of cyanogen gas to be C_2N_2. Hydrocyanic acid has been studied

under the head of acids, (page 100), and each metallic cyanide of interest to the physician or the pharmacist will be noticed in its appropriate place. Cyanogen unites with non-metallic elements, chlorine, iodine, etc., to form compounds which are of great interest to chemists, but of no use in medicine. Part second of this work, "Organic Chemistry," may be considered a continuation of the compounds of carbon.

METALS.

The metals as a class possess certain properties which the non-metals do not possess, or at least very slightly. It must be remembered that the division of the elements into the sections, metals and non-metals, is purely for convenience, and that the differences between certain individuals of the two classes is sometimes very slight. Chemists do not even agree as to the positions certain elements should occupy. For example some class hydrogen with the non-metals, others place it among the metals. Some class arsenic with the metals, others with non-metals. The generic property of metals seems to reside in metallic lustre, and in the fact that their oxides are generally basic. The oxides of the non-metals, upon the contrary, are acid formers, and as a rule non-metals have no metallic-like lustre.†

The metals are generally arranged in certain groups dependent upon a resemblance in chemical properties. We shall not, therefore, take them alphabetically. In our work we are expected to notice such compounds as are of interest, directly or indirectly, from a therapeutical point, but we shall from necessity, consider many oxides not used in medicine.

† One of the modifications of phosphorus is called *metallic phosphorus*, from its luster and resemblance to metals. Arsenic certainly appears like a metal. The higher *metals* give rise to acid-forming oxides.

ALKALI METALS.

POTASSIUM, SODIUM.
LITHIUM, RUBIDIUM.
CÆSIUM.

These are the metals that produce the *alkalies*, a term that was at first applied indiscriminately to the hydrates of potassium and of sodium. The alkalies were considered simple substances until Davy isolated the metals which form their bases. The compounds of potassium and sodium are extensively employed in medicine and the arts. Lithium is sometimes used in medicine, but the other members of this group of metals are of no practical importance.

POTASSIUM (*Kalium*).

Symbol, K. Atomic weight, 39.04.

Davy discovered this metal and named it. He decomposed caustic potash by means of the galvanic battery. Afterward Gay-Lussac and Thénard obtained it by heating caustic potash and iron turnings to whiteness in a gun barrel; the oxygen of the caustic potash uniting with the iron to form oxide of iron, thus liberating the hydrogen and the potassium, which escaped in the form of vapor, and from this the potassium was condensed by passing the mixture into a cold receiver. This process was improved upon by Curauda who obtained potassium by heating caustic potash with charcoal, and by this latter method it is made at present, or by heating carbonate of potassium with charcoal, thus:—

$$\underset{\text{Carbonate of potassium}}{K_2CO_3} + \underset{\text{Carbon}}{C_3} = \underset{\text{Potassium.}}{K_2} + \underset{\text{Carbonic oxide.}}{3CO.}$$

Properties.—Potassium is a white metal of specific gravity 0.875 at 13° C. (55.4 F.); soft like wax at 15° C. (59° F.); brittle at 0° C. (32° F.), and melts at 62.5° C. (144.5° F.) Potassium has a strong affinity for oxygen and a cleanly cut surface immediately tarnishes and turns

white; when thrown upon the surface of water it liberates hydrogen and forms hydrate or hydroxide of potassium, thus:—

$$\underset{\text{Potassium.}}{K_2} + \underset{\text{Water.}}{2(HOH)} = \underset{\text{Hydroxide of potassium.}}{2(KOH)} + \underset{\text{Hydrogen.}}{H_2}.$$

The heat liberated during the reaction inflames the escaping hydrogen, which burns with a violet color consequent upon the presence of the vapor of potassium, (pure hydrogen burns blue). In consequence of this affinity of potassium for oxygen it is necessary to preserve the metal beneath a liquid free from oxygen, and benzine is used in commerce for this purpose. Potassium reaches us from the manufacturer in the form of pellets ranging in size from a number one shot to that of a small marble.

Potassium Hydroxide.—Formula, KOH. Molecular weight, 56.00.

Synonyms.—Hydrate of potassium. Caustic potash. Fixed alkali. Potassa. Potash. Vegetable alkali.

This is derived from water by replacement of one-half the hydrogen of two molecules of water with one molecule of potassium.

It is obtained in an impure form by lixivating the ashes of forest trees and evaporating the "lye" to dryness. This is *potash* and as thus made is mixed with impurities such as salts of aluminum, potassium carbonate and chloride, potassium silicate, etc., and is unfit for medical use or chemical experiments. Commercial caustic potash is in the form of white sticks, and is made by boiling a solution of carbonate of potassium, or of bicarbonate of potassium, with milk of lime. The potassium is displaced from one molecule of carbonate of potassium (where that salt is used), by the calcium of one molecule of hydroxide of calcium, thus:—

$$\underset{\text{Carbonate of potassium.}}{K_2CO_3} + \underset{\text{Hydroxide of calcium.}}{CaO_2H_2} = \underset{\text{Hydroxide of potassium}}{2KOH} + \underset{\text{Carbonate of calcium.}}{CaCO_3}$$

The carbonate of calcium is permitted to subside, and the clear solution decanted and quickly evaporated until it fuses, when it is cast into moulds. This still contains

other salts of potassium, such as sulphates and chlorides, but is pure enough for all practical purposes. It may be produced pure for analytical purposes, by decomposing pure potassium nitrate with metallic copper, at a red heat, and by other processes unnecessary to name here.

Liquor potassæ of the Pharmacopœia, is made by boiling solution of bicarbonate of potassium with milk of lime, and decanting the clear solution, or by dissolving one troy ounce of caustic potash in a pint of distilled water. It is liable to contain the impurities of the potash salt. It should have the specific gravity 1.065, and contain "five and eight-tenths per cent. of hydrate of potassa," U. S. P.

Properties.—Hydroxide of potassium *pure* is white, hard, breaks with a fibrous fracture, melts to a clear liquid at red heat, volatilizes at higher temperatures, and the vapor decomposes at a white heat into its constituent elements. It dissolves in about one-third its weight of water, and from this crystallizes upon cooling in the form of a *hydrate*, $KOH+2H_2O$. It dissolves perfectly in alcohol. Commercial caustic potash is white, but does not dissolve completely in alcohol, owing to the presence of sulphates and alumina. Caustic potash possesses an acrid taste, even in dilute solution. It is a strong base, and from it the salts of potassium are prepared. It absorbs both moisture and carbonic acid from the atmosphere, and thus either deliquesces (moisture) or becomes covered with a white incrustation (carbonate of potassium) when exposed. It is a powerful cautery, and destroys animal and vegetable tissues, and thus is a *destructive poison* when swallowed.

Treatment.—Dilute acids neutralize it, and the juice of lemons, sour oranges, green gooseberries, ripe currents, and vinegar may be administered freely. In conjunction, use mucilaginous and demulcent drinks, such as gruel, milk, etc., or olive oil. If a large dose of the caustic potash has been swallowed, the tissues of the stomach will be actually destroyed before the antidote can be given, and the patient will suffer excruciating agony until relieved by death.

Potassium Acetate.—Formula, $KC_2H_3O_2$. Molecular weight, 97.9.

Synonyms.—Acetate of potassa. Acetate of potash. Potassii acetas.

Experiment.—Dissolve about sixty grains of carbonate of potassium in an ounce of distilled water in a porcelain dish, and gradually stir into the solution acetic acid until the liquid acquires a distinct acid reaction and ceases to effervesce. Evaporate carefully over a flame until a white mass remains; this is acetate of potassium. The reaction is expressed as follows:—

$$\underset{\text{Carbonate of potassium.}}{K_2CO_3} + \underset{\text{Acetic acid.}}{2HC_2H_3O_2} = \underset{\text{Acetate of potassium.}}{2KC_2H_3O_2} + \underset{\text{Water}}{H_2O} + \underset{\text{Carbon dioxide.}}{CO_2}$$

Acetate of potassium is extensively used in medicine. It is in the form of white glistening masses, of a foliated texture, and is made by evaporating a solution of acetate of potassium until a skim forms upon the surface, which is removed constantly until the dish is empty. From time to time a little acetic acid must be added to replace that which is lost by evaporation. Sometimes it is made by decomposition of acetate of lead, by means of carbonate of potassium, and in this case the salt may contain lead as an impurity. Sometimes the concentration of the solution is performed partially in copper vessels, and under such circumstances, especially if the solution is allowed to cool within the vessel, *copper* is liable to contaminate the result.

Properties and Tests.—Acetate of potassium is very hygroscopic, and quickly liquefies upon exposure to the atmosphere. It should not yield a precipitate when sulphide of hydrogen is passed through a solution of it, and the solution should remain transparent upon treatment with solution of chloride of barium.

Potassium Bicarbonate.—Formula, $KHCO_3$. Molecular weight, 99.89.

Synonyms.—Bicarbonate of potash. Acid carbonate of potassium. Potassæ bicarbonas. Hydrogen potassium carbonate. Potassii bicarbonas.

Experiment.—Make a saturated solution of carbonate of potassium by agitating one ounce of carbonate of potassium, with half an ounce of cold water; after the excess of carbonate has settled, decant the clear solution, pass into this solution a current of carbon dioxide, and in a short time small crystals will appear upon the sides and bottom of the container. This reaction is explained as follows:—

$$\underset{\text{Carbonate of potassium.}}{K_2CO_3} + \underset{\text{Water.}}{H_2O} + \underset{\text{Carbon dioxide.}}{CO_2} = \underset{\text{Bicarbonate of potassium.}}{2KHCO_3}$$

Bicarbonate of potassium may be made by exposing wet carbonate of potassium to an atmosphere containing much carbon dioxide, as the cellar of a brewery where fermentation is ensuing, or the cellar of a vinegar factory.

Salæratus is an impure bicarbonate of potassium, at one time much used for culinary purposes. It is made by passing the fumes of burning charcoal over shelves covered with layers of pearlash.

Properties.—Bicarbonate of potassium is met with in commerce in the form of crystals, which are transparent and permanent in cool air, but which become moist in damp, warm weather. It has an alkaline taste, and an alkaline reaction, notwithstanding it is an *acid* salt. It dissolves in about four times its weight of water at ordinary temperatures, and parts with a considerable proportion of carbonic acid, when its solution in water is boiled. Bicarbonate of potassium is more apt to be pure than is carbonate of potassium, owing to the fact that the impurities do not crystallize.

Potassium Bichromate.—Formula, $K_2Cr_2O_7$. Molecular weight, 294.6.

Synonyms.—Bichromate of potassium. Bichromate of potash. Red chromate of potash. Dichromate of potassium. Potassii bichromas.

The chemistry of this article is noticed under *chromium.* To prepare it in practice, chrome iron ore ($FeOCr_2O_3$) is roasted, and mixed with lime and potassium carbonate,

and the mixture heated in an oxidizing flame. The result is then extracted with water, and the neutral chromate converted into bichromate by means of sulphuric acid, thus:—

$$\underset{\text{Potassium chromate.}}{2K_2CrO_4} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Potassium bichromate.}}{K_2Cr_2O_7} + \underset{\text{Potassium sulphate.}}{K_2SO_4} + \underset{\text{Water.}}{H_2O}.$$

The formation of bichromate of potassium may be illustrated by the following experiment:—

Dissolve five or ten grains of chromium trioxide CrO_3 (officinal chromic acid), in half an ounce of liquor potassæ. A lemon yellow solution will result, dependent upon the production of neutral potassium chromate, thus:—$CrO_3 + 2KOH = K_2CrO_4 + H_2O$. Add to this solution a little sulphuric acid, an orange color results, dependent upon the formation of bichromate of potassium, as explained by the equation which precedes the last. If a very great excess of sulphuric acid is now added, the bichromate of potassium is decomposed and chromic acid is formed.

Bichromate of potassium is used in medicine mainly for the preparation of *chromic acid* (see page 93). It is extensively used in the arts, especially as a dye-stuff, in the preparation of chrome yellows and to make ink. Bichromate of potassium forms large garnet-red crystals, freely soluble in water. The dust is very irritating when inhaled, and the salt is a corrosive poison.

Treatment of Poisoning.—Emetics, albuminous and mucilaginous drinks (Taylor), carbonate of magnesium or chalk mixed with milk.

Potassium Bitartrate —Formula, $KHC_4H_4O_6$. Molecular weight, 187.68.

Synonyms.—Bitartrate of potassium. Cream of tartar. Acid tartrate of potassium. Potassii bitartras.

Bitartrate of potassium exists in the juices of many fruits (see tartaric acid) and is obtained from the juice of the grape after fermentation, which causes it to crystallize upon the sides of the casks. This is impure and known as crude tartar or argols.

Purification.—Dissolve argols in hot water, and filter

the solution while hot through clay or animal charcoal. Repeat the operation if necessary, and lastly crystallize.

Properties.—Bitartrate of potassium, when pure, is in the form of transparent crystals, soluble in fifteen parts of boiling water, but requiring 200 parts of cold water. It does not contain water of crystallization.

Impurities and Tests.—This salt is usually sold upon the market in a powdered condition, and under the name of cream of tartar. It is liable to contain moisture, and from five to fifteen per cent. of tartrate of calcium, for tartrate of calcium will dissolve in a solution of cream of tartar. Too often, however, alum, powdered silica, and terra alba (a white earth), etc., are mixed with it. Such substances as are insoluble will remain undissolved upon boiling the suspected cream of tartar with water. A current of sulphide of hydrogen will precipitate copper and lead from a solution of the salt. Starch may be shown by digesting the sample in ammonia water, boiling the residue (if any) with water, and then adding a little tincture of iodine, a blue color denoting starch. If adulterated with *alum* it is more soluble in cold water than if pure, and upon fusing it the residue will not completely dissolve if alum is present.

Potassium Bromide. — Formula, KBr. Molecular weight, 118 79.

Synonyms.—Potassii bromidum. Bromide of potash. Bromide of potassium.

Experiment —Pour a few ounces of solution of bromide of iron into an evaporating dish, and gradually add with constant stirring, dilute solution of carbonate of potassium, until the liquid loses the property of changing blue litmus paper to red, or until a precipitate ceases to form. Bring the mixture to a boil, then when cold filter it and wash the contents of the filtering paper with a little distilled water. Mix the filtrates, and evaporate until a pellicle begins to form upon the surface, then cool, when crystals of bromide of potassium will form. By the foregoing operation carbonate of iron and bromide of potassium are formed, thus :—

$$\underset{\text{Bromide of Iron.}}{FeBr_2} + \underset{\text{Carbonate of potassium.}}{K_2CO_3} = \underset{\text{Carbonate of iron.}}{FeCO_3} + \underset{\text{Bromide of potassium.}}{2KBr}$$

It may be made by other methods, but the above for experiment is easiest and most simple.

Properties.—Bromide of potassium crystallizes in cubes, sometimes transparent, again porcelain like. These forms may be produced at will from the same solution, and as the demand is greatest for the opaque crystals, manufacturers are inclined to make it of that style. If there is an excess of carbonate of potassium in the solution, the opaque bromide forms most readily. Bromide of potassium is permanent in the air, and if a specimen deliquesce an excess of carbonate of potassium is present. It is soluble in its weight of boiling water, and in less than twice its weight of water at 0° C. (32° F.). If a crystal be dipped into hydrochloric acid it should not turn yellow (absence of bromate of potassium).

Potassii Carbonas—Formula, K_2CO_3. Molecular weight, 137.93.

Synonyms.—Carbonate of potash. Carbonate of potassium. Salt of tartar.

Carbonate of potassium is made from the potash of wood ashes, or from beet roots (in Europe), or from the Stassfurt beds of chloride of potassium. Once it was made by calcining cream of tartar, whereby it derived the name *salt of tartar*. This last process yields a very pure article. It may be made pure by calcining bicarbonate of potassium, thus:—

$$\underset{\text{Bicarbonate of potassium.}}{2KHCO_3} = \underset{\text{Water.}}{H_2O} + \underset{\text{Carbon dioxide.}}{CO_2} + \underset{\text{Carbonate of potassium.}}{K_2CO_3}$$

Experiment.—Dissolve a little hydroxide of potassium (caustic potash) in four times its weight of water, then pass carbon dioxide into the solution. Carbonate of potassium will be formed according to the following equation:—

$$\underset{\text{Hydroxide of potassium.}}{2KOH} + \underset{\text{Carbon dioxide.}}{CO_2} = \underset{\text{Carbonate of potassium.}}{K_2CO_3} + \underset{\text{Water.}}{H_2O}$$

It is usually made in this country from potash, and thus is liable to contain many impurities.

Properties.—Carbonate of potassium is in the form of a white, granular, semi-crystalline coarse powder, deliquescent in the atmosphere and soluble in less than its weight of water at ordinary temperatures. It crystallizes from saturated solution, forming $2K_2CO_3+3H_2O$.

Pearlash is an impure carbonate of potassium and is made by dissolving potash in water, decanting the solution, and evaporating it to dryness. It contains caustic potash, and various salts such as sulphate and chloride of potassium, etc.

Carbonate of potassium is poisonous. The antidotes and treatment are the same as for hydroxide of potassium (caustic potash).

Potassium Chlorate.—Formula, $KClO_3$. Molecular weight, 122.29.

Synonyms.—Potassii chloras. Chlorate of potash.

Chlorate of potassium may be made by several processes. It is manufactured in quantities from chlorinated lime and chloride of potassium. When chlorinated lime is mixed with water and evaporated the *calcium hypochlorite* of the solution is decomposed, calcium chlorate and calcium chloride being formed, thus:—

$$\underset{\text{Hypochlorite of calcium.}}{3Ca(OCl)_2} = \underset{\text{Chlorate of calcium.}}{Ca(ClO_3)_2} + \underset{\text{Chloride of calcium.}}{2CaCl_2}.$$

If chloride of potassium (KCl) be added previous to evaporation, double decomposition ensues between the chlorate of calcium and the chloride of potassium, the result being chlorate of potassium instead of chlorate of calcium of the previous equation, thus:—

$$\underset{\text{Chlorate of calcium.}}{Ca(O_3Cl)_2} + \underset{\text{Chloride of potassium.}}{2KCl} = \underset{\text{Chloride of calcium.}}{CaCl_2} + \underset{\text{Chlorate of potassium.}}{2KClO_3}.$$

By straining the solution and evaporating it to crystallization the chlorate of potassium is separated from the very soluble chloride of calcium, and it may be purified by solution in and recrystallizing from hot water.

Chlorate of potassium is only made in large quantities where chlorine gas is a waste product, and it is not advisable to experiment with its production. It crystallizes in transparent tables of the monoclinic system which melt at 334° C., and which begin to decompose, evolving oxygen at 352° C. One hundred parts of water dissolve 3.3 parts of the salt at 0° C. (32° F.) and at 104.8° C. (about 221° F.) dissolve 60.2 parts.†

Chlorate of potassium must never be triturated in a mortar with organic bodies which have an affinity for oxygen, for chlorate of potassium readily gives up its oxygen, and serious explosions have resulted. Mixtures of sugar, tannin, sulphur, etc., should never be triturated with potassium chlorate, the safe method being to powder the salts separately, then mix them upon a paper by means of a spatula. Chlorate of potassium has a saline, cooling taste.

Impurities and Tests.—Dissolve a portion of the chlorate of potassium in hot water, and add a few drops of nitric acid, then divide this solution into two parts, and to one portion add solution of nitrate of silver; if *chloride of potassium* be present, a white precipitate will be produced, soluble in ammonia water.

To the other portion add solution of nitrate of barium; a white precipitate will denote *sulphate of potassium*. There are a few other likely impurities.

Potassium Citrate.—Formula, $K_3C_6H_5O_7$. Molecular weight, 305 66.

Synonyms.—Citrate of potassium. Citrate of potash. Potassæ citras. Potassii citras.

This is a salt formed by double decomposition, between citric acid and carbonate of potassium, thus:—

† Physicians often overlook the fact that chlorate of potassium is but slightly soluble in cold water, and prescribe (to be made into a solution) a very much larger amount of the salt than can be dissolved. Of course if heat be employed to effect the solution the excess of chlorate will separate when the liquid cools. Water at ordinary temperature (60° F.) will dissolve about one-sixteenth its weight of chlorate of potassium.

$$3K_2CO_3 + 2H_3C_6H_5O_7 = 2K_3C_6H_5O_7 + 3H_2O + 3CO_2$$

Carbonate of potassium. | Citric acid. | Citrate of potassium. | Water. | Carbon dioxide.

Experiment.—Dissolve a few drachms of carbonate of potassium in an evaporating dish, by means of a couple of ounces of water, then carefully stir into it a solution of citric acid until effervescence ceases, then evaporate with constant stirring to dryness. Citrate of potassium is a white granular powder. It is very soluble in water, and has a pleasant saline taste, but is not extensively used in medicine.

Potassium Cyanide.—Formula, KCNorKCy. Molecular weight, 65.02.

Synonyms.—Cyanide of potassium. Cyanide of potash. Potassa cyanide Potassii cyanidum.

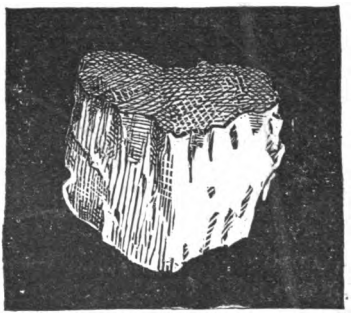

FIG 53. Potassium Cyanide.

This is made pure by passing the vapor of hydrocyanic acid into an alcoholic solution of hydroxide of potassium (caustic potash), when cyanide of potassium will separate as a white powder, the reaction being expressed as follows:

$$HCN + KOH = KCN + HOH.$$

Hydrocyanic acid. | Hydroxide of potassium. | Cyanide of potassium. | Water.

The above yields a chemically pure salt, if *pure* caustic potash is employed. In commerce, cyanide of potassium is made by melting ferro-cyanide of potassium (yellow prussiate of potash) with carbonate of potassium; the

result being cyanide of potassium (KCN), cyanate of potassium (KOCN), carbon dioxide (CO_2) and iron, thus:

$$2K_4FeCy_6 + 2K_2CO_3 = 10KCN + 2KOCN + 2CO_2 + Fe_2$$

Ferrocyanide of potassium. Carbonate of potassium. Cyanide of potassium. Cyanate of potassium. Carbon dioxide. Iron.

Cyanide of potassium, prepared as above, is contaminated with *cyanate* of potassium.

Properties.—This salt, as found in the market, is in white lumps, resembling in appearance chloride of calcium, or the white substance found inside the cocoanut. It has a strong bitter-almond or peach-pit odor, owing to the fact that it constantly evolves hydrocyanic acid, especially when exposed to an atmosphere which contains carbon dioxide. It is very soluble in water, is a powerful reducing agent, and hence is employed in photography and electro-plating. It is one of the most *violent* of poisons, and must never be confounded with *prussiate* of potassium (ferrocyanide of potassium), which is sometimes prescribed by physicians. Bear in mind the distinctions.

Prussiate of potash is in form of yellow crystals; the crystals are odorless; it is not classed with poisons.

Cyanide of potassium is in form of white lumps; evolves the odor of bitter almonds; is a deadly poison.

Treatment.—Solution of sulphate of iron may be given; this will (theoretically) form Prussian blue. The treatment for hydrocyanic acid may be followed. If even a moderate dose is swallowed, however, there is little hope of recovery, for the soluble salts of hydrocyanic acid are as poisonous as the acid alone.

Potassium and Sodium Tartrate. — Formula, $KNaC_4H_4O_6\,4H_2O$. Molecular weight, 281.51.

Synonyms.—Sodii et potassii tartras. Potassæ et sodæ tartras. Tartrate of potash and soda. Rochelle salt. Potassii et sodii tartras.

This, as its name indicates, is a salt of which one-half of the basic hydrogen of the tartaric acid molecule is replaced by potassium, the other half being replaced by sodium, thus:—

Tartrate of hydrogen (tartaric acid), $HHC_4H_4O_6$.
Acid tartrate of potassium (cream of tartar), $KHC_4H_4O_6$.
Tartrate of potassium and sodium (Rochelle salt) $KNaC_4H_4O_6$.

Experiment.—Dissolve twelve parts of carbonate of sodium in eighty-four parts of hot water, and stir into the solution sixteen parts of bitartrate of potassium. When effervescence ceases, filter the solution and evaporate it until a pellicle begins to form, then place the solution in a cool place that crystals may form. These are tartrate of potassium and sodium. The common name *Rochelle salt*, was derived from Rochelle, where it was discovered. This salt is made by manufacturers upon a large scale, and is usually sold in the form of powder. It has a saline taste, and dissolves in twice its weight of water at about 60° F. The reaction of the preceding process is expressed by the following formula:—

$$\underset{\text{Carbonate of sodium.}}{Na_2CO_3} + \underset{\text{Bitartrate of potassuim.}}{2KHC_4H_4O_6} = \underset{\text{Tartrate of sodium and potassium.}}{2KNaC_4H_4O_6} + \underset{\text{Water.}}{H_2O} + \underset{\text{Carbon dioxide.}}{CO_2}.$$

Potassium Ferrocyanide. — Formula, $K_4Fe(CN)_6$. Molecular weight, 367.94. Crystallized, $K_4Fe(CN)_6,3H_2O$.

Synonyms.—Ferrocyanide of potassium. Yellow prussiate of potash. [Prussiate of potash.] Potassii ferrocyanidum.

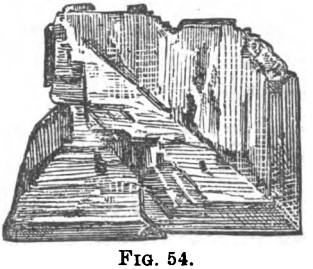

FIG. 54.

This salt results when a solution of a ferrous salt is mixed with an excess of cyanide of potassium, but is made in practice by heating potash, iron-filings and animal matter, such as horn, dried blood, leather scraps, etc.

An intricate reaction results, cyanogen (C_2N_2) being formed from the decomposing organic substances which unite with the potassium to form cyanide of potassium, while the liberated sulphur and the sulphur of the potash produces potassium and iron sulphides. The sulphide of iron and the cyanide of potassium are mutually decomposed to form ferrocyanide of potassium and other products according (R. and S.) to the following equation :—

$$13KCN + Fe_2S_3 = 2K_4Fe(CN)_6 + 2K_2S + KSCN.$$

Cyanide of potassium. | Se-qui-sulphide of iron. | Ferrocyanide of potassium. | Potassium mono-sulphide. | Potassium sulpho-cyanate

This mixture is dissolved and purified, and the ferrocyanide of potassium crystallized. It is not practical to make this salt in an experimental way.

Ferrocyanide of potassium is insoluble in alcohol, but dissolves in four parts of cold water and in two parts of hot water. It crystallizes in large yellow masses of quadratic pyramids, has a saline sweetish taste, and completely gives off its water of crystallization at 100° C. (212° F.) It is not dangerous, being sometimes prescribed in large doses. Care must be taken that it be not confounded with *cyanide of potassium* (see page 178), which is very poisonous.

Potassium Hypophosphite,—Formula, KPH_2O_2. Molecular weight, 103.92.

Synonyms.—Potassæ hypophosphis. Hypophosphite of potassium. Hypophosphite of potash. Potassii hypophosphis.

Hypophosphite of potassium may be made by decomposing hypophosphorous acid and carbonate of potassium, and afterward evaporating the solution carefully to dryness, thus :—

$$K_2CO_3 + 2HPH_2O_2 = 2KPH_2O_2 + H_2CO_3.$$

Carbonate of potassium. | Hypophosphorous acid. | Hypophosphite of potassium. | Carbonic acid.

It may also be made by double decomposition between solutions of carbonate of potassium and hypophosphite of calcium, the result being hypophosphite of potassium and precipitated carbonate of calcium.

Hypophosphite of potassium is very soluble in water and diluted alcohol. It is not likely to contain impurities which are soluble in these menstruums unless it be the carbonate of potassium. If this substance be present the addition of dilute sulphuric acid to the solution will cause effervescence.

Potassium Iodide.—Formula, KI. Molecular weight, 165.57.

Synonyms.—Iodide of potassium. Iodide of potash. Potassii iodidum.

This salt may be made exactly according to the formula of our experiment for making bromide of potassium, (page 174), excepting that the solution of bromide of iron of that experiment is to be replaced by solution of iodide of iron. The reaction is expressed as follows:—

$$\underset{\text{Iodide of Iron.}}{FeI_2} + \underset{\text{Carbonate of potassium.}}{K_2CO_3} = \underset{\text{Iodide of potassium.}}{2KI} + \underset{\text{Carbonate of iron.}}{FeCO_3}.$$

It may also be made by adding iodine to solution of caustic potash until in slight excess, then evaporating to dryness and igniting the residue with charcoal, and afterward dissolving the iodide of potassium and crystallizing it.

Properties.—Iodide of potassium crystallizes either in transparent or porcelain cubes, it is odorless, and possessed of a saline bitter taste. It is not altered by dry air, but turns yellowish in damp air. If it contain carbonate of potassium the crystals become moist upon exposure and will eventually dissolve in the water so attracted. The crystals dissolve in three-fourths their weight of cold water and in one-half their weight of boiling water, and dissolve in from five to seven times their weight of alcohol.

Impurities and Tests.—*Iodate* of potassium is shown by placing a drop of hydrochloric acid upon the crystal, when if iodate be present a brown color will immediately follow; or boil a few grains of starch with some water in a test tube and dissolve in the solution a crystal of the suspected iodide of potassium, then add a few drops of a strong solution of tartaric acid; if *iodate* of potassium be present a violet color will immediately appear. Ad-

mixtures such as sulphate and nitrate of potassium may be shown by dissolving iodide of potassium in its weight of water, then mixing the solution with twelve parts of absolute alcohol, a white precipitate indicating such impurities. *Chlorides* or *bromides* may be detected as follows: Dissolve a small portion of the salt in water and add nitrate of silver solution in excess; agitate the precipitate with ammonia water in excess, filter and then add to the filtrate an excess of nitric acid; a white precipitate denotes *chlorine*, (of course the nitric acid is to be free from chlorine). Bromine may be distinguished from chlorine (Hoffman) by washing the foregoing precipitate upon a filtering paper until the washings cease to redden blue litmus paper, then placing the precipitate in a test tube and agitating with chlorine water; *bromine*, if present, will be liberated and will color the solution *yellow*. Bromide of potassium is the most likely adulterant of this salt.

Potassium Nitrate. — Formula, KNO_3. Molecular weight, 100.93.

Synonyms.—Nitrate of potassium. Nitrate of potash. Saltpetre. Nitre.[Sal nitri. Sal petræ (Geber).] Potassii nitras.

Nitrate of potassium is found in the soil of many countries, in the juices of many plants, in the soil of underground caverns. It is formed by the decomposition of organic matter when such is in contact with potash salts, and in this manner the soil replenishes itself when depleted.

Experiment.—Dissolve an ounce of carbonate of potassium in four fluid ounces of water in a capacious porcelain dish, then gradually stir into it nitric acid until effervescense ceases. Evaporate the solution to the bulk of two fluid ounces, and cool slowly to the temperature of ice water. The nitrate of potassium will form in transparent crystals, which may be drained, and dried upon filtering paper. The reaction is expressed as follows:—

$$\underset{\text{Carbonate of potassium.}}{K_2CO_3} + \underset{\text{Nitric acid.}}{2HNO_3} = \underset{\text{Nitrate of potassium.}}{2KNO_3} + \underset{\text{Carbonic acid.}}{H_2CO_3}.$$

The nitrate of potassium of commerce is supplied largely from the soil of India, and recently by decomposition from the chloride of potassium of the Stassfurt mines. The so-called Chili saltpetre is nitrate of sodium.

Properties.—Nitrate of potassium usually crystallizes in rhombic prisms, but from a few drops of solution by spontaneous evaporation the form is rhombohedral. The large crystals of the market are often hollow, and partly filled with mother liquor, which renders this form more impure than small crystals. It melts at 339° C. (642.2° F.) forming a corlorless liquid, which, if moulded and cooled into balls, is known upon the market as *sal prunella;* at a high temperature the melted salt gives off oxygen and nitrogen. It is insoluble in alcohol, but very soluble in water. 100 parts of water at 0° C. (32° F.) dissolves 13.3 parts; 100 parts of water at 100° C. (212° F.) dissolves 247 parts.

Impurities and Tests.—The most likely adulterant of nitrate of potassium is *nitrate of sodium.* If it be present in any great proportion, the crystals will become moist upon exposure to a moderately damp atmosphere. If the solution of nitrate of potassium acidulated with nitric acid, gives a white precipitate with nitrate of barium solution, *sulphates* are present. If it gives a white precipitate with nitrate of silver solution, *chlorides* are present. Little of the nitrate of potassium upon the market is pure, and the presence of a small proportion of the foregoing substances is not a serious matter.

† **Potassium Permanganate.**—$KMnO_4$, or $K_2Mn_2O_8$. Molecular weight ($KMnO_4$), 157.68.

Synonyms.—Permanganate of potassium. Permanganate of potash. Potassii permanganas.

This is made experimentally by dissolving ten parts of caustic potash in the smallest amount of water practicable, and adding to it a mixture composed of seven parts of potassium chlorate, and eight parts of dioxide of manganese, and evaporating to dryness. The mass is digested with hot water, and the deep green liquid filtered through asbestos, when upon standing crystals of permanganate of

potassium will be deposited. Permanganate of potassium is a salt of the very unstable *permanganic* acid ($HMnO_4$). It is in crystals of dark bronze or steel-blue color, of the specific gravity 2.7; soluble in fifteen or sixteen parts of cold water, the solution being dark purple. The solution of this salt is decolorized by most organic substances, and by inorganic reducing agents, such as sulphurous acid, and metallic subsalts. It ignites some oxydizable bodies, therefore mixtures of potassium permanganate and organic substances must be made with caution.

Potassium Sulphate.—Formula, K_2SO_4. Molecular weight, 173.90.

Synonyms.—Sulphate of potassium. Sulphate of potash. Potassii sulphas.

This was obtained at a very early day as a residue of the manufacture of nitric acid. It is found in the lavas of Vesuvius and in the Stassfurt deposits. It may be made by neutralizing sulphuric acid, which has previously been diluted with ten parts of water, with hydroxide of potassium (caustic potash), and evaporating the solution to crystallization; the reaction being expressed as follows:

$$\underset{\text{Sulphuric acid.}}{H_2SO_4} + \underset{\text{Hydroxide of potassium.}}{2KOH} = \underset{\text{Sulphate of potassium.}}{K_2SO_4} + \underset{\text{Water.}}{2H_2O}$$

It may be also made by the following—

Experiment.—Mix one fluid ounce of sulphuric acid with nine fluid ounces of distilled water, and into this mixture gradually stir carbonate of potassium until the acid is saturated and effervescence ceases. Evaporate the solution until a pellicle forms upon the surface, then allow it to cool, when crystals of sulphate of potassium will form, which may be dried upon filtering paper. The reaction is expressed according to the equation:—

$$\underset{\text{Sulphuric acid.}}{H_2SO_4} + \underset{\text{Carbonate of potassium.}}{K_2CO_3} = \underset{\text{Sulphate of potassium.}}{K_2SO_4} + \underset{\text{Carbonic ac.d.}}{H_2CO_3}$$

Sulphate of potassium crystallizes in small rhombic pyramids, having the specific gravity 2.648. It is soluble in ten parts of cold water; a boiling solution of the salt

contains about one-fourth its weight of the salt. It is insoluble in absolute alcohol.

Acid Sulphate of Potassium, $KHSO_4$, is obtained when nitrate of potassium is decomposed with sulphuric acid in the manufacture of nitric acid (see page 105). This is sometimes called *bisulphate of potassium,* but is seldom used in medicine.

Potassium Disulphate, $K_2S_2O_7$, is obtained by heating sulphuric acid with sulphate of potassium, thus:

$$H_2SO_4 + K_2SO_4 = K_2S_2O_7 + H_2O.$$

This is of no practical value.

Potassium Sulphide.

Synonyms—Sulphuret of potassium. Potassii sulphidum. Sulphide of potassium. Liver of sulphur. Hepar sulphuris. Potassii sulphuretum.

This compound as prepared by the various pharmacopœias, is a mixture of different sulphides of potassium. It is made (U. S. P.) by fusing one troy ounce of sulphur with two troy ounces of carbonate of potassium, and pouring the fused mass upon a cold slab of marble to cool, and immediately transferring to glass stoppered bottles. The product is influenced by the temperature employed, so that it is variable in appearance.

Tests for Potassium.—Compounds of potassium, in the absence of sodium compounds, impart a violet color to the non-luminous blow-pipe flame. If sodium compounds be present, a blue cobalt glass placed between the eye and the flame, will absorb the yellow color of sodium, and transmit the potassium violet.

2d. Add an excess of solution of the chloride of platinum ($PtCl_4$), to the solution of a potassium salt, and evaporate the mixture to dryness. Digest the residue in alcohol, filter it and dry, and weigh the precipitate, which is the double salt K_2PtCl_6, from which the proportion of potassium may be determined. Tartaric acid also precipitates the solution of a potassium salt.

SODIUM.

Symbol, Na. Atomic weight, 22.99.

Sodium was discovered by Davy in 1807, being produced by decomposition of hydroxide of sodium (caustic soda) by means of the galvanic current, after the manner in which he obtained potassium. Chloride of sodium and carbonate of sodium have been known from the earliest periods of history, the term *nitrum* being undoubtedly applied to carbonate of sodium at an early day and afterward affixed to saltpetre. Metallic sodium is now made by igniting a mixture of carbonate of sodium and charcoal, upon the principle of the preparation of potassium. Sodium is much more easily made than potassium. Sodium is a silver-white metal, but quickly oxidizes upon exposure to the atmosphere. When heated in the atmosphere combustion ensues, the flame being yellow, and the product of the combustion being the monoxide (Na_2O), and the dioxide (Na_2O_2) which is sometimes called the peroxide. Sodium melts at 95.6° C. (203.9° F.) and volatilizes at a red heat. It is an excellent conductor of heat and electricity. When thrown into cold water it swims upon its surface with a hissing noise, decomposing the water and evolving from it hydrogen gas. If the water be hot the hydrogen is ignited, burning with a yellow flame, the yellow color being produced by the vapor of sodium. Sodium reaches market in the form of sticks that have been cut from sheets half an inch or so in thickness. It is preserved in benzine. Sometimes these sticks are several inches in length. It will be remembered that potassium occurs in the form of pellets, and that potassium will ignite the liberated hydrogen when thrown on *cold* water, and that the flame of this burning hydrogen is *pink* instead of yellow.

Sodium Hydroxide.—Formula, NaOH Molecular weight, 39.95

Synonyms.—Sodium hydrate. Caustic soda.

This compound results when sodium is thrown upon water, and may be produced by adding water to the mon-

oxide of the hydroxide of sodium. Caustic soda, however, is manufactured in the arts by boiling lime with carbonate of sodium and subsequently evaporating the solution and fusing as follows:—

$$\underset{\text{Carbonate of sodium.}}{Na_2CO_3} + \underset{\text{Quicklime.}}{CaO} + \underset{\text{Water.}}{H_2O} = \underset{\text{Hydroxide of sodium.}}{2NaOH} + \underset{\text{Carbonate of calcium.}}{CaCO_3}$$

Caustic soda is white, brittle, melts below a red heat and has the specific gravity 2.13. Large crystals are deposited when a solution of the specific gravity 1.365 is cooled to—8° C. (17.6° F.) These crystals have the composition $2NaOH+7H_2O$. Large amounts of caustic soda are used in making soap, and the so-called concentrated lye of commerce is in reality caustic soda. In any form caustic soda is a *corrosive poison* if taken into the stomach, destroying the tissues with which it comes into contact and occasioning intense suffering. *Treatment*, same as for caustic potash. (see page 170).

Sodium Acetate. — Formula, $NaC_2H_3O_2$. Molecular weight, 81.85.

Synonyms.—Sodæ acetas. Sodii acetas. Acetate of sodium.

This salt is the product of the double decomposition of hydroxide of sodium and acetic acid, or of carbonate of sodium and acetic acid. It may be made by the following

Experiment.—Add crystallized carbonate of sodium in form of powder to a few ounces of acetic acid until effervescence ceases. Then evaporate the solution until a pellicle forms, and cool the solution; crystals of acetate of sodium will form. The reaction is expressed as follows:

$$\underset{\text{Sodium carbonate.}}{Na_2CO_3} + \underset{\text{Acetic acid.}}{2HC_2H_3O_2} = \underset{\text{Sodium acetate.}}{2NaC_2H_3O_2} + \underset{\text{Carbonic acid.}}{H_2CO_3}.$$

The crystallized acetate of sodium contains three molecules of water, being represented by $NaC_2H_3O_2+3H_2O$. It dissolves in less than half its weight of boiling water, and in about three and one-half parts of cold water.

Sodium Arseniate.—Formula of fresh crystals (not officinal), $Na_2HAsO_4,12H_2O$. Molecular weight, 401.24. Formula of exposed salt, $Na_2HAsO_4.7H_2O$. Molecular weight, 311 44.

Synonyms.—Sodii arsenias. Arseniate of sodium.

This is a poisonous compound made from sodium and arsenic, and is sometimes used in medicine. It is of great importance in calico printing, as it is the only reliable substitute for an offensive substance used in clearing the cloth after mordanting. It is made by heating a mixture of arsenic (arsenious acid of commerce, arsenious anhydrid, As_2O_3) and carbonate of sodium and nitrate of sodium, to redness. The mass when cold is dissolved in hot water and crystallized by cooling. There are several sodium arseniates and they bear a close resemblance to the sodium phosphates, but are of little importance in medicine. It might be supposed that the foregoing formula would produce *arsenite* of sodium. This is not the case, however, for by the roasting of the mixture nitrous anhydrid (N_2O_3) and carbon dioxide (CO_2) escape and *pyroarseniate* of sodium ($Na_4As_2O_7$) remains. When the pyroarseniate of sodium is dissolved in water it is converted into ortho-arseniate of sodium by combining with one molecule of water, thus: $Na_4As_2O_7+H_2O=2Na_2HAsO_4$, and when this crystallizes and is dried the *arseniate* of sodium remains, ($Na_2HAsO_4.7H_2O$).

Arsenite of Sodium.—This is formed when arsenious anhydrid is dissolved in solution of caustic soda. It is very unstable and is not used in medicine. The treatment for poisoning by arseniate of sodium is the same as that for arsenious acid.

Sodium Benzoate.—Formula, $NaC_7H_5O_2$. Molecular weight, 143.70.

Synonyms.—Benzoate of sodium. Benzoate of soda. Sodii benzoas. Sodæ benzoas.

Experiment.—Mix one ounce of benzoic acid† with

† Use benzoic acid which is made from gum benzoin instead of that from urine.

eight fluid ounces of water in a capacious porcelain evaporating dish, and warm it to the temperature of 180° F., and then gradually stir powdered carbonate of sodium into the mixture until effervescence ceases and the benzoic acid is dissolved. Now filter the solution and evaporate the filtrate to dryness. The reaction is expressed as follows:—

$$\underset{\text{Benzoic acid.}}{2HC_7H_5O_2} + \underset{\text{Sodium carbonate.}}{Na_2CO_3} = \underset{\text{Sodium benzoate.}}{2NaC_7H_5O_2} + \underset{\text{Carbonic acid.}}{H_2CO_3}.$$

Benzoate of sodium is a white salt, odorless or possessing a faint odor of benzoin, (*not* of urine unless made from the acid derived from urine,) and is very soluble in water.

Sodium Bicarbonate.—Formula, $NaHCO_3$. Molecular weight, 83.84.

Synonyms.—Sodii bicarbonas venalis. Sodæ bicarbonas. Bicarbonate of soda. [Baking soda. Soda.] Hydrogen sodium carbonate. Acid carbonate of sodium.

This salt is manufactured on a very large scale and is never made by the pharmacist. It may be prepared by the following process as an experiment. Dry three parts of crystallized carbonate of sodium, by exposure to a moderate heat in an iron vessel. Then mix this in a mortar with two parts of crystallized carbonate of sodium and place the mixture in a box, in a very thin layer, and pass carbonic acid through the box until the gas ceases to be absorbed by the powder. The reaction will be represented by the formula:—

$$\underset{\substack{\text{Sodium}\\\text{carbonate.}}}{Na_2CO_3} + \underset{\text{Water.}}{H_2O} + \underset{\substack{\text{Carbon}\\\text{dioxide.}}}{CO_2} = \underset{\substack{\text{Sodium}\\\text{bicarbonate.}}}{2NaHCO_3}$$

In the above the water is derived from the crystallized carbonate of sodium ($Na_2CO_3.10H_2O$.) If the entire amount of carbonate of sodium employed were of the crystallized salt, the water liberated would make a mush of the mixture, but by driving off the water of crystallization from a part, by the previous application of heat, the trouble is avoided.

Bicarbonate of sodium is found in market as a white powder, odorless, of a saline cooling taste, soluble in water though less so than the normal carbonate. It always contains carbonate of sodium, but the impurity is not serious. The carbonate may be removed by washing the bicarbonate with small portions of water until the washings cease to yield a precipitate with sulphate of magnesium. After it is dry, carbonate of sodium will be found to have formed again. When bicarbonate of sodium in solution is boiled carbon dioxide escapes. When this solution is permitted to cool, crystals separate which have the composition $Na_2CO_3+2HNaCO_3+3H_2O$, and are thus a combination of the two carbonates of sodium. A natural salt of this combination is found in many countries and is known as *Trona* or *Urao*.

Sodium Bisulphite.—Formula, $NaHSO_3$. Molecular weight, 103.85.

Synonyms.—Hydrogen sodium sulphite. Acid sulphite of sodium. Bisulphite of soda.

Bisulphite of sodium is made by passing sulphur dioxide into a concentrated solution of carbonate of sodium and afterward drying the crystals which separate. Sodium sulphite first forms, then the reaction is expressed:

$$\underset{\text{Sodium sulphite.}}{Na_2SO_3} + \underset{\text{Sulphur dioxide.}}{SO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Sodium bisulphite.}}{2NaHSO_3}.$$

Sodium bisulphite has the odor and taste of burning sulphur and an acid reaction. This salt must not be confused with

Sodium Disulphite $Na_2S_2O_5$ which results, according to Muspratt, from passing sulphur dioxide into a hot saturated solution of sodium carbonate.

Sodium Bromide.—Formula, NaBr. Molecular weight, 102.74.

Synonym.—Bromide of sodium.

This is prepared precisely after the directions for making bromide of potassium, carbonate of sodium being substituted for the carbonate of potassium of that process. If the crystals of bromide of sodium be deposited from

a hot concentrated solution, they are anhydrous and cubical. If crystallized from an ordinary solution, by evaporation at ordinary temperatures, monoclinic prisms form containing two molecules of water, ($NaBr2H_2O$). Bromide of sodium is soluble in less than its weight of hot water, and in one and one-fourth parts of cold water.

Sodium Carbonate.—Formula, Na_2CO_3. Molecular weight, 105.83. Crystallized—formula, $Na_2CO_3 + 10H_2O$. Molecular weight, 285.43.

Synonyms.—Sodii carbonas. Carbonate of sodium. Carbonate of soda. Sal-soda. Washing soda. Normal carbonate of soda. [Soda-ash (anhydrous carbonate of sodium).]

Carbonate of sodium is produced in immense quantities in the alkali manufacture. It is found in the waters of many lakes, in the ashes of plants from the sea shore and from marine plants, and is found in the soils of some countries. When sulphate of sodium is ignited with chalk and coal dust carbonate of sodium results. It is unnecessary to give an experiment here for making this salt as it is never made even upon an ordinary scale, and may be purchased for a few cents per pound.

Properties.—Carbonate of sodium crystallizes from a concentrated solution which is slowly allowed to cool, in large transparent crystals of the composition $Na_2CO_3 + 10H_2O$. Sometimes these crystals weigh many pounds. They have the specific gravity, 1.45, and are soluble in two parts of water and evolve heat in the act of solution. Carbonate of sodium has an alkaline taste and reaction, and effloresces upon exposure to a dry atmosphere, falling into a white powder.

The commercial crystallized carbonate usually contains sodium chloride and sulphate, and even other impurities. These may be separated by recrystallization.

There are carbonates of sodium containing different amounts of water of crystallization, but such are not used in medicine.

Sodium Chloride.—Formula, NaCl. Molecular weight, 58.36.

Synonyms.—Chloride of sodium. Salt. Common salt.

Chloride of sodium is the most widely and abundantly distributed of salts. It occurs in the form of beds in many places, in the waters of many springs, in various soils, in the waters of some inland lakes and in the ocean. It is one of the most important of substances, and yet, from its very abundance, requires little attention from us, and is never made from its component elements.

Properties.—Chloride of sodium usually crystallizes in the form of cubes, yet rock-salt is found as octohedrons and in forms intermediate between the cube and octohedron. It has the specific gravity at 0° C. (32° F.) of 2.16; it melts at 776° C. and volatilizes, unchanged, at a little higher degree. It is slightly more soluble in warm than in cold water. At 0° C. (32° F.) water dissolves 35.5 parts of chloride of sodium, at 5° C. (41° F) it dissolves 35.63 parts, and at 100° C. (212° F.) dissolves 39.16 parts. Chloride of sodium sometimes contains, as impurities, sulphate of sodium and the chlorides of magnesium, iron, and calcium. These are unimportant as regards its use in pharmacy unless they exist in larger amount than is possible with the purified table salt. If chloride of calcium be present to any appreciable extent, the salt will become moist upon exposure to the atmosphere.

Liquor Sodæ Chlorinatæ, U. S. P.

Synonyms.—Solution of chlorinated soda. Labarraque's solution.

This is made as follows:—

Dissolve twenty-four troy ounces of carbonate of sodium in three pints of hot water, and mix it with a solution of chlorinated lime previously prepared by triturating twelve troy ounces of chlorinated lime with nine pints of water. Allow the mixture to settle, decant the clear solution and wash the residue with water sufficient to make the liquid measure eight pints.

Solution of chlorinated soda is colorless and evolves the disagreeable odor of chlorine. It decomposes upon ex-

posure, and owing to this fact and the variable quality of the chlorinated lime of the market of which it is prepared, is liable to be of uncertain strength. It is an excellent disinfectant.

Sodium Hypophosphite.—Formula, $NaPH_2O_2+H_2O$. Molecular weight, 105.83.

Synonyms.—Hypophosphite of sodium. Hypophosphite of soda.

This salt may be made by the process employed in making hypophosphite of potassium (see page 181) substituting in this case carbonate of sodium for the carbonate of potassium of that formula, thus:—

$$Na_2CO_3+2HPH_2O_2=2NaPH_2O_2+H_2CO_3.$$

Hypophosphite of sodium is generally made in practice by double decomposition between hypophosphite of calcium and carbonate of sodium, the solution of hypophosphite of sodium being afterward filtered from the precipitated carbonate of calcium, after which it (solution) is evaporated carefully to dryness.

Properties.—Hypophosphite of sodium may be obtained in pearly crystals, but that of commerce is in the form of a white granular powder, or in warty masses. It will keep in a dry atmosphere, but deliquesces in moist air. It is very soluble in water, has a salty, alkaline taste, and dissolves in alcohol. When strongly heated it evolves phosphoretted hydrogen which inflames, the residuum being a mixture of the meta and the pyrophosphates of sodium.

Sodium Hyposulphite. — Formula, $NaHSO_4$. Molecular weight, 119.81.

This salt is made by the action of metallic zinc upon aqueous solution of sulphurous acid, or the solution of acid sulphite of sodium. It is *not* used in medicine and is not the salt generally known under the name of Hyposulphite of sodium. That which follows is—

Commercial Hyposulphite of Sodium.—Formula, $Na_2S_2O_3+5H_2O_3$. Molecular weight, 247.62.

Synonyms—Thiosulphate of sodium (new name). Hyposulphite of soda (old name).

This salt is made in immense quantities for use in the arts by boiling sulphite of sodium with sulphur, or by boiling sulphur and soda-lye together and then passing sulphur dioxide into the solution until it is colorless. In either case the solution is evaporated to crystallization.

Hyposulphite of sodium crystallizes as large transparent crystals resembling Glauber's salt. They are possessed of a cooling saline taste, are permanent in the air, have the specific gravity 1.672, dissolve freely in water, are insoluble in alcohol, and are odorless. The aqueous solution of it decomposes and deposits sulphur. It is employed in photography, and in dyeing, and in the manufacture of paper.

Sodium Iodide.—Formula, NaI. Molecular weight, 149 52.

Synonyms.—Sodii iodidum. Iodide of sodium.

This salt may be made by the process recommended upon page 174 for the preparation of bromide of potassium. In this instance solution of iodide of iron is to be first made from iodine and iron, and this by double decomposition with solution of carbonate of sodium will yield solution of iodide of sodium and precipitated carbonate of iron. The solution of iodide of sodium is to be obtained by filtration from the carbonate of iron, and afterward evaporated to dryness, and removed to well stoppered bottles.

Properties.—Iodide of sodium is a white crystalline powder, and is deliquescent. It crystallizes in anhydrous cubes from hot concentrated solution, but by spontaneous evaporation monoclinic prisms which contain two molecules of water are formed. Iodide of sodium dissolves in less than half its weight of hot water and in two-thirds its weight of cold water. It has a saline taste.

Sodium Nitrate.—Formula, $NaNO_3$. Molecular weight, 84.88.

Synonyms.—Sodii nitras. Nitrate of sodium. Cubic saltpetre. Chili saltpetre. Cubic nitre.

Nitrate of sodium may be made in an experimental manner by adding carbonate of sodium to diluted nitric acid until effervescence ceases. The solution is then to be evaporated to crystallization, and permitted to cool, that crystals may form. Nitrate of sodium is found in large deposits in the rainless districts of Peru and Bolivia, associated with sulphate and chloride of sodium and sulphate of calcium. It is exported from these countries in immense quantities in a more or less purified condition, and for the purpose of making nitric acid is used instead of the more expensive nitrate of potassium. Sodium nitrate has the specific gravity 2.26, is very soluble in water and deliquesces in moist air. It crystallizes transparent and the crystals are obtuse rhombohedrons.

Sodium Phosphate. — Formula, $Na_2HPO_4 12H_2O$. Molecular weight, 357.30.

Synonyms.—Sodii phosphas. Phosphate of sodium. Hydrogen disodium orthophosphate.

Phosphoric acid is tribasic (H_3PO_4) and is therefore capable of yielding three different phosphates :—

1. Trisodium orthophosphate, $Na_3PO_4+12H_2O$.
2. Hydrogen disodium orthophosphate, $Na_2HPO_4+12H_2O$
3. Dihydrogen sodium orthophosphate, $NaH_2PO_4+H_2O$.

The second of these we find in market under the name phosphate of sodium, or phosphate of soda, and as the others are of no use in medicine we will only consider the :—

Hydrogen Disodium, Orthophosphate, or ordinary phosphate of sodium. This may be made by neutralizing orthophosphoric acid with carbonate of sodium, thus—

$$\underset{\text{Sodium carbonate.}}{Na_2CO_3} + \underset{\text{Phosphoric acid.}}{H_3PO_4} = \underset{\text{Phosphate of sodium.}}{Na_2HPO_4} + \underset{\text{Carbonic acid.}}{H_2CO_3}.$$

It is made practically from bone-ash, but the process is tedious and scarcely necessary to inculcate into a work of this description.

Properties.—Phosphate of sodium crystallizes in large transparent monoclinic prisms which contain twelve molecules of water of crystallization. They give off a portion of this water upon exposure to dry air, and effloresce. They have the specific gravity of 1.525 and dissolve freely in water, but are insoluble in alcohol. With solution of nitrate of silver a solution of this salt produces a *yellow* precipitate. When calcined they part with their water of crystallization and form what is known as—

Normal Sodium Pyrophosphate, or *pyrophosphate of sodium.* $Na_4P_2O_7$.

This is a glassy mass, soluble in water, and the solution has a strongly alkaline reaction. When the solution is boiled with an acid the salt changes to the ordinary phosphate (Stromeyer). When a solution of pyrophosphate of sodium is evaporated to crystallization, the crystals which form contain ten molecules of water ($Na_4P_2O_7 + 10H_2O$). Pyrophosphate of sodium is used to make the so-called pyrophosphate of iron.

Sodium Salicylate.—Formula, $Na_2C_7H_4O_3$. Molecular weight, 181.65.

Synonyms.—Salicylate of sodium. Salicylate of soda. This salt is made easily by the following—

Experiment.—Mix one part of pure salicylic acid with eight parts of water, in a porcelain dish, and heat to the temperature of 80° C. (176° F.), then gradually stir into this a strong solution of carbonate of sodium until it ceases to produce effervescence, and a clear solution is produced. Add now a few grains of salicylic acid, and if effervescence follow continue the addition of salicylic acid until it occasions no farther effervescence. Then filter the solution by means of a funnel the exit of which is plugged with a wad of cotton. Lastly, evaporate the solution to dryness over a water bath. During the addition of the sodium carbonate, and afterward while evaporating the

solution, stir constantly with a porcelain or wooden spatula. The reaction is explained as follows:—

$$H_2C_7H_4O_3 + Na_2CO_3 = Na_2C_7H_4O_3 + H_2CO_3.$$

Salicylic acid. Sodium carbonate. Sodium salicylate. Carbonic acid.

If an iron spatula be employed to stir the solution, or an ordinary filtering paper to filter it, a red color will result.

Properties.—Salicylate of sodium has sometimes a pinkish appearance. It is very soluble in water, and strikes a red color with ferric salts. It should be white.

Sodium Silicate.—Formula, $Na_2Si_4O_9$. Molecular weight, 301.62,

Synonyms.—Sodii silicas. Silicate of sodium. Silicate of soda. Soluble glass.

Silicate of sodium is made upon the large scale by dissolving flint under pressure in a hot solution of caustic soda, or by heating white sand, soda-ash and charcoal together.

Properties.—Soluble glass of the market usually contains an excess of caustic soda. It is generally sold in the form of a thick solution, odorless, and of a yellowish or amber color. It is but little used in medicine, although it is a powerful agent in arresting putrefaction. It is employed as a mucilage, for making soap, for painting, for making artificial stone, and in other branches of the arts. When nearly neutral it is used to adulterate certain kinds of fancy candies. The silicate of sodium of commerce is made in salt works where caustic soda is abundant.

Sodium Sulphate.—Formula, Na_2SO_4. (anhydrous) Molecular weight, 141 80. Crystallized,—$Na_2SO_4 + 10H_2O$. Molecular weight, 321.4.

Synonyms.—Sodii sulphas. Sulphate of sodium. Sulphate of soda. [Sodium deca-hydrate.] Glauber's salt. (Salt-cake, for the anhydrous).

Sulphate of sodium is found native (anhydrous), in the waters of certain mineral springs, in the waters of the ocean and salt lakes, and is an incidental product of many chemical operations. This salt was first mentioned by Glauber, and from this fact derived the common name,

Glauber's salt. It is very cheap and common at the present day. It may easily be made as an experiment by saturating a solution of carbonate of sodium with dilute sulphuric acid, thus:—

$$\underset{\text{Sulphuric acid.}}{H_2SO_4} + \underset{\text{Sodium carbonate.}}{Na_2CO_3} = \underset{\text{Sodium sulphate}}{Na_2SO_4} + \underset{\text{Carbonic acid.}}{H_2CO_3}.$$

The salt may be obtained in the form of crystals from the foregoing solution by evaporating it to the proper consistence and allowing the solution to cool.

Properties.—Glauber's salt crystallizes from cold aqueous solution in colorless mono-clinic prisms, which effloresce in dry air. They melt in their water of crystallization at 33° C. (91.4° F.) and become anhydrous upon exposure to a temperature of less than 100° C. (212° F.) One hundred parts of water dissolve fifty-five parts of Glauber's salt at 34° C. (93.2° F.) and above this temperature, as the heat is increased, there is a constant decrease of the amount of salt dissolved until at 103.5° C. (218.3° F.) only 42.2 parts are held in solution. This is explained by the fact that above 34° C. the deca hydrate ($Na_2SO_4 + 10H_2O$) begins to decompose, giving rise to the anhydrous salt (Na_2SO_4) of much less solubility.

Sodium Sulphite. — Formula, Na_2SO_3. Molecular weight, 125.84. Crystallized, $Na_2SO_3 7H_2O$. Molecular weight, 251.56.

Synonyms.—Sodii sulphis. Sulphite of sodium. Sulphite of soda. Normal sodium sulphite.

Sulphite of sodium is made by passing sulphur dioxide into a solution of carbonate of sodium until the solution will change the color of blue litmus paper to red, then evaporating it to crystallization. The reaction is as follows:—

$$\underset{\text{Sodium carbonate.}}{Na_2CO_3} + \underset{\text{Sulphur dioxide.}}{SO_2} = \underset{\text{Sodium sulphite.}}{Na_2SO_3} + \underset{\text{Carbon dioxide.}}{CO_2}.$$

Sulphite of sodium may also be made by adding solution of sulphurous acid to solution of carbonate of sodium until an acid reaction results, and afterward evaporating the solution to crystallization, thus:—

$$\underset{\text{Sodium carbonate.}}{Na_2CO_3} + \underset{\text{Sulphurous acid.}}{H_2SO_3} = \underset{\text{Sodium sulphite.}}{Na_2SO_3} + \underset{\text{Carbonic acid.}}{H_2CO_3}.$$

Still another process is to pass sulphur dioxide into solution of carbonate of sodium until the gas ceases to be absorbed. Then dissolve in the solution an amount of sodium carbonate equal in weight to the sodium carbonate originally employed. This process is explained by the fact that when solution of carbonate of sodium is *saturated* with sulphur dioxide, *bisulphite* of sodium is formed, thus:—

$$\underset{\text{Sodium carbonate.}}{Na_2CO_3} + \underset{\text{Sulphur dioxide.}}{2SO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Sodium bisulphite.}}{2NaHSO_3} + \underset{\text{Carbon dioxide.}}{CO_2}.$$

And when an equal amount of carbonate of sodium is added to this solution of bisulphite of sodium, carbonic acid is displaced and normal sulphite of sodium results as follows:—

$$\underset{\text{Sodium carbonate.}}{Na_2CO_3} + \underset{\text{Sodium bisulphite.}}{2NaHSO_3} = \underset{\text{Sodium sulphite.}}{2Na_2SO_3} + \underset{\text{Carbonic acid.}}{H_2CO_3}.$$

Properties.—Sulphite of sodium is deposited in transparent crystals of the monoclinic system. They have a strong sulphurous taste, and in solution react alkaline towards litmus paper. They are soluble in their weight of boiling water and in about five parts of cold water. They absorb oxygen from the air and change to *sulphate* of sodium. This property is also possessed by solution of sulphite of sodium, therefore it is essential that the bottle containing it should be securely corked. When heated it loses its water of crystallization. The dried salt is prepared as a medicine, but is less soluble than the crystallized.

Tests for Sodium.—Sodium compounds impart a yellow color to a jet of burning hydrogen. The most delicate test for sodium is spectrum analysis.

If the solution of a sodium salt be mixed with an excess of platinum tetrachloride solution, a precipitate is produced (Na_2PtCl_6) which dissolves readily in alcohol. Advantage is taken of this fact in the separation of sodium and potassium compounds, for the potassium chloroplatinate thus produced is insoluble in alcohol.

LITHIUM.

Symbol, Li. Atomic Weight, 7.01.

Lithium (as *lithia*) was discovered in the year 1817. It derived its name from *litheios*, a word which signifies "stony." The salts of lithium are widely distributed throughout the animal and vegetable kingdoms, but in very small quantities. The waters of certain mineral springs contain lithium salts, notably the water from a spring in the Wheal Clifford mine in Cornwall, England. It is present in most soils, being derived from the disintegrating rocks. From the soils it is taken by vegetables and from the vegetable kingdom is carried to animals.

Properties.—Lithium, the metal, was first prepared in experimental quantities by Bunsen and Matthiessen in the year 1855, although Davy is said to have obtained traces of the metal by electrolysis. Lithium is a silver-like metal, and tarnishes by exposure to the atmosphere, though less readily than potassium and sodium. It is remarkable for being the lightest of the known solid elements, having the specific gravity of from 0.5891 to 0.5983, (water the standard). Consequently it will float upon water and even upon benzine. It is harder than potassium and sodium and softer than lead. Melts at 180° C. (356° F.) and when it is heated somewhat above this point in the air it burns with a white flame. Nitric acid acts upon it with violence. It burns if heated in chlorine, bromine vapor or iodine vapor, forming compounds of the respective elements. When thrown upon water it floats and rapidly oxidizes, but does not fuse. Lithium is not used in medicine or the arts, but a few of the salts of lithium are used in medicine.

Lithium Bromide.—Formula, LiBr. Molecular weight, 86.76.

Synonyms.—Lithii bromidum. Bromide of lithium.

This salt is most easily made by the following process. Place a few fluid ounces of freshly prepared hydrobromic acid in a capacious evaporating basin, and gradually stir

into it carbonate of lithium until effervescence ceases and the carbonate is no farther dissolved; then filter the solution and evaporate it to dryness. Lastly preserve the dry salt in well closed vials. The foregoing reaction is explained by the following equation:—

$$\underset{\text{Lithium Carbonate.}}{Li_2CO_3} + \underset{\text{Hydrobromic acid.}}{2HBr} = \underset{\text{Lithium bromide.}}{2LiBr} + \underset{\text{Carbonic acid.}}{H_2CO_3}.$$

Bromide of lithium is a white powder, but may be obtained in crystals by evaporating a concentrated solution of it over sulphuric acid. This salt is characterized by the fact that it contains a larger proportion of bromine than any other solid salt into which bromine enters.

Lithium Benzoate.—Formula, $LiC_7H_5O_2$. Molecular weight, 127.72.

Experiment.—Mix in a capacious porcelain evaporating dish one part of benzoic acid made from gum benzoin with eight parts of water, and having warmed the mixture, gradually stir into it carbonate of lithium until effervescence ceases and the benzoic acid is dissolved. Now filter the solution and evaporate it over a water bath until a dry salt remains. Benzoate of lithium is white and odorless, or it may possess a slight agreeable odor of benzoin. It should never be employed when it exhales the odor of putrid urine, and such may be the case if the salt is impure from having been made from benzoic acid which was produced from urine. If solution of benzoate of lithium be evaporated until a pellicle forms and is then cooled, it will form colorless satin-like crystals.

Lithium Carbonate.—Formula, Li_2CO_3. Molecular weight, 73.87.

Synonyms.—Lithii carbonas. Carbonate of lithium. Carbonate of lithia. Normal lithium carbonate.

This salt may be made by pouring a solution of chloride of lithium into a solution of carbonate of ammonium in ammonia water, and then heating the mixture as long as a precipitate forms. This precipitate is to be dried by exposure to the atmosphere.

Carbonate of lithium is a white powder, odorless and very slightly soluble in both hot and cold water. It dissolves perfectly in most acid solutions forming the respective salts. Oxide of iron is sometimes present as an impurity in carbonate of lithium. Carbonate of lithium is of interest from the fact that it is employed in pharmacy to prepare all the salts of lithium that are used in medicine.

Bicarbonate of Lithium, or *hydrogen lithium carbonate,* $LiHCO_3$, is the form in which lithium is found in some mineral waters. It is soluble to the extent of five parts in one hundred of water, and it may be made artificially by spontaneously evaporating a saturated solution of carbonate of lithium until a crystalline deposit results, and afterward passing carbon dioxide through the mixture until the crystals dissolve again. This salt is unstable and upon evaporation of the solution, or exposure of it to the atmosphere, carbon dioxide escapes and the normal salt is deposited, thus:—

$$\underset{\text{Hydrogen lithium carbonate.}}{2LiHCO_3} = \underset{\text{Lithium carbonate.}}{Li_2CO_3} + \underset{\text{Carbonic acid.}}{CO_2 + H_2O}.$$

Lithium Citrate.—Formula, $Li_3C_6H_5O_7$. Molecular weight, 209.57.

Synonyms. Lithii citras. Citrate of lithium. Citrate of lithia.

Experiment.—Dissolve one part of citric acid in eight parts of water in a capacious evaporating dish, and having warmed the solution, gradually stir into it carbonate of lithium until effervescence ceases. Then filter the solution of citrate of lithium, return it to the dish, and evaporate it to dryness on a sand bath. The reaction by which citrate of lithium is prepared by the foregoing formula is expressed as follows:—

$$\underset{\substack{\text{Lithium}\\\text{carbonate.}}}{3Li_2CO_3} + \underset{\text{Citric acid.}}{2H_3C_6H_5O_7} = \underset{\substack{\text{Lithium}\\\text{citrate.}}}{2Li_3C_6H_5O_7} + \underset{\substack{\text{Carbon}\\\text{dioxide.}}}{3CO_2} + \underset{\text{Water.}}{3H_2O}.$$

Citrate of lithium is a white powder, and must be preserved in well closed vials, as it deliquesces in moist air. It is very soluble in water.

Lithium Iodide. — Formula, LiI. Molecular weight, 133.54.

Synonyms.—Lithii iodidum. Iodide of lithium.

This salt may be produced by the process recommended for making bromide of lithium, *hydriodic* acid being substituted for hydrobromic acid. The remarks we have applied to bromide of lithium may be also accepted for this salt, iodine being substituted for the word bromine.

Tests for Lithium Compounds.—Minute proportions of this element are detected by the spectroscope. The hydrogen flame is colored carmine-red by lithium compounds. Salts of lithium may be detected by this method. Strontium is liable to be confused with lithium, but the flame is colored of a purer scarlet by salts of lithium, than by salts of strontium. If a lithium flame be examined through an indigo prism, it appears carmine-red where the solution is thin and fainter in color, as the thicker portion of the prism is approached, at last being entirely overcome. Potassium under like conditions is blue where the prism is thin, changing to violet as it thickens, at the *center* becoming crimson-red. If an ordinary neutral solution of a lithium salt be boiled with solution of sodium phosphate (Na_2HPO_4), a crystalline precipitate of normal lithium phosphate (Li_3PO_4) results. In this manner lithium can be separated from potassium and sodium.

RUBIDIUM.

Symbol, Rb. Atomic weight, 85.2.

Bunsen discovered this element in the year 1861, by the aid of spectrum analysis. It exists in many minerals, saline springs and earths, and is taken up in minute amounts by the vegetable kingdom. Neither rubidium nor any of its salts is used in medicine.

CÆSIUM.

Symbol, Cs. Atomic weight, 132.5.

Cæsium is remarkable from the fact that it is the first element which was discovered by means of the spec-

troscope. Bunsen discovered it in the year 1860, just preceding his discovery of rubidium. Up to our day the pure element has not been obtained, although small globules of the metal appear at the negative pole when a current of electricity is passed through melted chloride of cæsium (electrolysis). These rise to the surface and burn brilliantly. Cæsium salts are found widely disseminated throughout the mineral kingdom and in many saline waters, but not in plants. None of the compounds of cæsium are used in medicine.

AMMONIUM SALTS.

When a small piece of moistened sal-ammoniac is placed within a platinum dish, and a little mercury is poured upon it, and the negative pole of a galvanic battery is attached to the mercury, and the positive pole to the platinum, the sal-ammoniac is decomposed. The mercury swells and forms what is known as *ammonium amalgam.* This may also be made by placing a globule of mercury the size of a pea in a test tube, and adding to it potassium in small pieces until a lump of potassium the size of the mercury has been used. Pour upon the amalgam a strong solution of chloride of ammonium. The globule at the bottom of the tube expands to a silvery mass (ammonium amalgam), until often the entire tube is filled, then with escape of ammonia and hydrogen it will suddenly shrink to its former size.

Upon these facts we find some chemists base their belief in the existence of a radical *ammonium,* which they suppose is similar to potassium and sodium in some respects, but so readily decomposed into simple elements as to debar isolation and thus forbid examination. Be this as it may, the theory of a radical ammonium (NH_4) is the basis of our present notation of the salts known as *ammonium salts,* and these salts are at this day considered to be derived from acids by replacement of their hydrogen by the radical NH_4 (ammonium). Thus as an example we may say that NH_4OH is hydroxide of ammonium, upon the same principle that we say KOH is hydroxide of

potassium. Hydroxide of ammonium, however, is not permanent and separates into a compound of the composition NH_3 (ammonia gas) and (H_2O) water. All endeavors to obtain the radical NH_4 have failed, if we may except the amalgam theory before mentioned. The compound (NH_4) is univalent, and replaces one atom of hydrogen or of any element that can replace an atom of hydrogen. Example:—

$$NH_4OH + HNO_3 = NH_4NO_3 + H_2O.$$
Ammonium hydroxide. Nitric acid. Ammonium nitrate. Water.

We may also argue that hydroxide of ammonium is a solution of ammonia gas in water, and in this case express our equation as follows:—

$$NH_3 + H_2O + HNO_3 = NH_4NO_3 + H_2O.$$
Ammonia gas. Water. Nitric acid. Ammonium nitrate. Water.

In either instance the result may be accepted as a compound of NH_4 with the nitric acid radical NO_3.†

If we accept the radical ammonium, the salts of it should be classed with those of potassium and sodium. Practically they are similar, and the potassium or sodium of a compound may be replaced by the combination NH_4. For this and other reasons we shall follow the first group of metals with the *ammonium salts.*

Ammonium Acetate.—Formula, $NH_4C_2H_3O_2$. Molecular weight, 76.87.

Acetate of ammonium is not permanent. It is made in solution by neutralizing acetic acid with carbonate of ammonium, or with aqua ammoniæ.

Acetate of ammonium dissolves in water, and is used by physicians under the name *Liquor Ammonii Acetatis,*

† These salts may also be called combinations of ammonia, NH_3, with an acid, thus:—

$$NH_3 + HCl = NH_3HCl.$$
Ammonia. Hydrochloric acid. Hydrochlorate of ammonia.

$$NH_3 + HNO_3 = NH_3HNO_3.$$
Ammonia. Nitric acid. Nitrate of ammonia.

According to this view ammonia is similar to organic bases (alkaloids), and thus forms salts by uniting bodily with acids.

U. S. P., or *Spiritus Mindereri.* The officinal process for making it is as follows: Take of diluted acetic acid two pints, and of carbonate of ammonium a sufficient quantity, add the carbonate gradually to the acid until the latter is neutralized, and then filter. When this preparation is dispensed, it should be freshly made.

Properties.—Solution of acetate of ammonium should be colorless, and should freely evolve gas bubbles when exposed to the atmosphere and stirred. It should not have a disagreeable smoky odor, and for this reason the acetic acid employed must be pure and free from empyreumatic products. If fresh it contains an abundance of carbonic acid, and has a sharp, pleasant taste. In consequence of the presence of the carbonic acid, blue litmus paper can not be used with satisfaction as a test for neutralization in its preparation. A better plan is to add a lump of carbonate of ammonium to the diluted acetic acid, and remove it immediately when effervescence ceases. Then add a little acetic acid to give the solution a pleasant acidity.

Ammonium Benzoate,—Formula, $NH_4C_7H_5O_2$. Molecular weight, 138.72.

Synonyms.—Ammonii benzoas. Benzoate of ammonium. Benzoate of ammonia.

Experiment.—Take of benzoic acid, from gum benzoin, two troy ounces; water of ammonia, three and a half fluid ounces, or a sufficient quantity; distilled water, four fluid ounces. Dissolve the acid in three and a half fluid ounces of the water of ammonia mixed with the distilled water; evaporate with a gentle heat, occasionally adding water of ammonia, if necessary, to maintain a slight excess of the alkali; then set aside to crystallize, and dry the crystals without heat. (U. S. P.)

Properties.—Benzoate of ammonium is in white crystals, which evolve ammonia when heated with solution of caustic potash or caustic soda. It is quite soluble in water, possesses a sweetish taste and should be either odorless or should simply exhale a slight aroma of gum benzoin. Under no condition should it be offensive to the sense of smell, or evolve the odor of stale urine.

Ammonium Bromide.—Formula, NH_4Br. Molecular weight, 97.76.

Synonyms—Ammonii bromidum. Bromide of ammonium.

Bromide of ammonium may be made by decomposing a solution of bromide of iron with a solution of carbonate of ammonium, filtering the solution of bromide of ammonium from the precipitated carbonate of iron, and evaporating the bromide of ammonium solution to dryness. It cannot be practically prepared by the officinal process of the U. S. P. where ammonia water is employed to decompose the bromide of iron solution, as ammonia will not accomplish the desired end, and the salt will then be discolored, or even red. Bromide of ammonium is most easily made for experimental purposes by neutralizing any given amount of hydrobromic acid with ammonia water, and then evaporating the solution to dryness. The reaction is as follows:—

$$\underset{\text{Hydrobromic acid.}}{HBr} + \underset{\text{Ammonium hydroxide.}}{NH_4OH} = \underset{\text{Ammonium bromide.}}{NH_4Br} + \underset{\text{Water.}}{H_2O}.$$

Bromide of ammonium is a white salt, very soluble in water and possessed of a sharp, salty taste. It leaves no residue when strongly heated upon a piece of platinum foil. Its solution is not precipitated by solution of barium chloride.

Commercial Ammonium Carbonate.—Formula, $N_3H_{11}C_2O_5 = H(NH_4)CO_3, NH_4CO_2NH_2$. Molecular weight, 156.77.

When a mixture of two parts of chalk and one part of chloride of ammonium is heated in a retort a white substance sublimes, known, after resublimation with a little water, as carbonate of ammonium. This is the old *salvolatile or salt of hartshorn.* It is not properly a true carbonate of ammonium, but a mixture of two substances, hydrogen ammonium carbonate (bicarbonate of ammonium) and ammonium carbamate.

Commercial ammonium carbonate is sometimes called sesquicarbonate of ammonia. It occurs in market as

translucent masses of a pungent ammoniacal odor and a sharp, caustic taste. According to Roscoe and Schorlemmer, it is composed of a molecule of bicarbonate of ammonium, $(NH_4)HCO_3$, and a molecule of carbamate of ammonium $NH_4CO_2NH_2$. Consequently its composition is represented by the empirical formula $N_3H_{11}C_2O_5$. Upon exposure the carbamate volatilizes, the lump turns white, opaque, loses its pungency, becomes of light weight, and finally only bicarbonate of ammonium remains. For this reason carbonate of ammonium should be kept in salt-mouth bottles, and should be discarded after it has undergone decomposition.‡ Carbonate of ammonium sometimes contains traces of iron salts which give it a pink tinge. It should be completely volatilized when heated to redness upon platinum foil, and if this is not the case, fixed impurities such as sulphate of calcium are present. Sometimes metallic lead is present in the form of small chips or shavings, such being scraped from the receiver in which the salt was condensed.

Hydrogen Ammonium Carbonate or Bicarbonate of Ammonium, $(NH_4)HCO_3$. This salt is sometimes found in crystalline form in Patagonian guano and in the purifiers of gas-works. It is the white substance which remains when commercial carbonate of ammonium is treated with alcohol, or is exposed to the atmosphere until it ceases to evolve the odor of ammonia. It is soluble in about eight parts of water at ordinary temperatures, is insoluble in alcohol and is not valued in medicine.

Normal Ammonium Carbonate, $(NH_4)_2CO_3+H_2O$. This salt is formed (R. and S.) when *ammonium carbamate* is dissolved in water as follows:—

$$\underset{\text{Ammonium carbamate.}}{NH_4CO_2NH_2} + \underset{\text{Water.}}{2H_2O} = \underset{\text{Ammonium carbonate.}}{(NH_4)_2CO_3+H_2O}.$$

It may be prepared otherwise, but the processes are of little general interest. It must be remembered that com-

‡The German Pharmacopœia recognizes *ammonium carbonicum pyro-oleosum* which is made by mixing thirty-two parts of finely powdered commercial carbonate of ammonium with ethereal animal oil *one* part.

mercial carbonate of ammonium contains *carbamate* of ammonium, and therefore, an aqueous solution of commercial carbonate of ammonium is to an extent a solution of the normal carbonate. This results from the combination of the carbamate of ammonium with water which, as is shown above, forms the normal carbonate of ammonium.

Carbamate of Ammonium, $NH_4CO_2NH_2$. This, as before stated, is a component of commercial carbonate of ammonium. It is formed by bringing dry carbon dioxide into contact with dry ammonia gas, or by passing the two gases into cold, absolute alcohol. In the reaction which follows one molecule of carbon dioxide combines with two molecules of ammonia as follows:—

$$\underset{\text{Carbon dioxide.}}{CO_2} + \underset{\text{Ammonia.}}{2NH_3} = \underset{\text{Ammonium carbamate.}}{NH_4CO_2NH_2}.$$

This salt is considered the valuable constituent of carbonate of ammonium; it has a pungent, ammoniacal odor, is a crystalline powder, and when heated to 140° C. (284° F.) it decomposes into urea and water, thus:—

$$\underset{\text{Ammonium carbamate.}}{NH_4CO_2NH_2} = \underset{\text{Urea.}}{CO(NH_2)_2} + \underset{\text{Water.}}{H_2O}.$$

Ammonium Chloride.—Formula, NH_4Cl. Molecular weight, 53.38.

Synonyms.—Ammonii chloridum. Chloride of ammonium. Muriate of ammonia. Hydrochlorate of ammonia. Sal-ammoniac.

This salt was introduced into Europe from Egypt, where it was made by burning camel's dung. It has been stated that it derived the original name, *sal-ammoniacum*, from the temple Jupiter Ammon, near the Libyan desert, but as this desert yields common salt, very likely the name was confused with the two salts.

Chloride of ammonium is found native in the neighborhood of volcanoes and in some specimens of guano. It may be easily made by neutralizing hydrochloric acid with ammonia water and afterward evaporating the solution to dryness, the reaction being expressed by the equation:

$$\underset{\text{Hydrochloric acid.}}{HCl} + \underset{\text{Hydroxide of ammonium.}}{NH_4OH} = \underset{\text{Ammonium chloride.}}{NH_4Cl} + \underset{\text{Water.}}{H_2O}.$$

In the arts chloride of ammonium is made by heating sulphate of ammonium with lime, as in the preparation of ammonia water, and passing the ammonia gas which is evolved into muriatic acid. This solution of chloride of ammonium is afterward evaporated to dryness and the dry salt volatilized and condensed. Or a mixture of chloride of sodium and sulphate of ammonium is heated, when by double decomposition chloride of ammonium is formed, and volatilizes while sulphate of sodium remains, thus:—

$$\underset{\text{Sodium chloride.}}{2NaCl} + \underset{\text{Ammonium sulphate.}}{(NH_4)_2SO_4} = \underset{\text{Sodium sulphate.}}{Na_2SO_4} + \underset{\text{Ammonium chloride.}}{2NH_4Cl}$$

Properties.—Chloride of ammonium is found in market in fibrous masses, very heavy and of a slight reddish or yellowish color (iron), particularly upon the surface. It is translucent and has a sharp saline taste. When pure it is colorless and odorless. It dissolves freely in water with much decrease of temperature, and is soluble to a considerable extent in alcohol. The most common impurity of the commercial salt is iron, which may be separated by dissolving the chloride of ammonium in water, adding a slight excess of ammonia water, filtering the product and afterward evaporating the filtrate to dryness. Formerly chloride of ammonium was employed in making ammonia water, but it is now replaced for this purpose by the cheaper sulphate of ammonium.

Ammonium Citrate.—Formula, $(NH_4)_3C_6H_5O_7$. Molecular weight, 242.57.

Synonyms.—Ammonii citras. Citrate of ammonia.

Experiment.—Dissolve three parts of citric acid in sixteen parts of water and add a sufficient amount of strong ammonia water to completely neutralize the acid. This is in accordance with the British Pharmacopœia, not being officinal in the U. S. P. Citrate of ammonium in the dry state is never used in medicine. The solution of citrate of ammonium is one of the ingredients of the scaled iron salt known as *pyrophosphate* of iron.

Ammonia.—Formula, NH_3. Molecular weight, 17.01.

Synonym.—Ammonia gas.

When a silent electrical discharge is passed through a mixture of nitrogen and hydrogen gases it is demonstrated (Donkin) that direct combination may ensue, and that a body of the above composition will in that case be formed. Ammonia has been known since a very early day. It exists in small amount in the atmosphere as a carbonate, and is found in combination with nitrous and nitric acids in rain water, and as chloride of ammonium it is obtained from several natural sources. In the arts it is made by heating an alkali with either the sulphate or the chloride of ammonium. Generally sulphate of ammonium and lime are employed on account of the cheapness of these salts. Ammonia gas in an impure form is derived almost exclusively from the organic kingdom for the purpose of making the salts of ammonium, and when these salts are purified, as we have already said, they are in turn decomposed in the practical production of ammonia.

At one time ammonia was made from urine. Then, by the distillation of hoofs, horns, bones, cartilage, etc.; but at present it is prepared from the liquor known as *ammoniacal liquor*, which is derived from the gasworks of our cities. When coal is distilled, the nitrogen present (about 2 per cent.) unites with hydrogen to form ammonia, which passes over with the coal gas. This gas is passed into water and the ammonia, together with some ammonium salts remains in solution. This "ammoniacal liquor" is distilled with lime and yields very impure *ammonia*.†

Properties.—Ammonia gas is an irrespirable, colorless gas, of alkaline reaction, becoming liquid at —40° C. (—40° F.)freezing to white crystals at —75° C. (—103° F.) It is very soluble in water, forming—

† See Ammonium Sulphate.

Ammonium Hydroxide. — Formula, $NH_4OH = NH_3H_2O$. Molecular Weight, 34.97.

Synonyms.—Ammonia. Water of ammonia. Hydrate of ammonium. [Spirit of hartshorn. Spiritus volatilis salis ammoniaci.] Liquor ammoniæ. Solution of ammonia. Aqua ammoniæ.

Ammonia gas will dissolve in water to a great extent, much heat being liberated by the condensation. Inasmuch as it is far more soluble in cold water than in warm, the receiver must be cooled by means of ice or running water while the gas is passing into it. The solution of ammonia gas in water decreases the specific gravity of the liquid. The U. S. P. recognizes two strengths as follows.

Aqua Ammoniæ, of specific gravity 0.960, one hundred grains of which neutralize thirty of officinal sulphuric acid.

Aqua Ammoniæ Fortior (stronger water of Ammonia), of the specific gravity 0.900. This contains twenty-six per cent. by weight of ammonia gas.

Ammonia water of whatever strength should not effervesce on the addition of dilute acids, and when neutralized with pure nitric acid should not yield precipitates with solutions of either nitrate of silver, carbonate of ammonium, or chloride of barium. It must be remembered that ammonia water parts readily with ammonia gas, most especially during warm weather, and therefore, great care must be exercised that the container be well stopped.

Spirit of Ammonia.—Under this name a preparation is sometimes used which consists of solution of ammonia gas in alcohol. This doubtless is sometimes confused with ammonia water, but generally in commerce where *spirit of ammonia* is mentioned, the solution of the gas in alcohol is recognized. Spirit of ammonia is officinal.

Ammonium Iodide.—Formula, NH_4I. Molecular, weight, 144.54.

Synonyms.—Ammonii iodidum. Iodide of ammonium. Iodide of ammonia.

This salt is prepared according to the U. S. P. by the double decomposition of iodide of potassium and sulphate

of ammonium. It is liable, as thus prepared, to be contaminated with sulphate of potassium. It is best made by the following experiment:—

Into an evaporating dish place any convenient amount of hydriodic acid, and then stir into it of water of ammonia a quantity sufficient to render the solution of alkaline reaction. Evaporate it quickly to dryness, and preserve the salt in well stoppered vials. The reaction is explained as follows:—

$$\underset{\text{Hydriodic acid.}}{HI} + \underset{\text{Ammonium hydroxide.}}{NH_4OH} = \underset{\text{Ammonium iodide.}}{NH_4I} + \underset{\text{Water.}}{H_2O}.$$

Iodide of ammonium may also be prepared by dissolving iodine in hydro-sulphide of ammonium, boiling the solution and filtering to separate sulphur, and then evaporating the filtrate to dryness.

Properties.—Iodide of ammonium is a white salt, of a sharp, saline taste, and very soluble in both water and alcohol. It becomes moist in damp air, and turns yellow from liberation of iodine. This salt should be kept in small, well stoppered vials.

Ammonium di-iodidum, NH_4I_2, is made by adding to a concentrated solution of nitrate of ammonium made alkaline with caustic potash, a little iodine. The salt separates as a black liquid. It is not used in medicine.

Ammonium Nitrate.—Formula, NH_4NO_3. Molecular weight, 79.90.

Synonyms.—Ammonii nitras. Nitrate of ammonium. Nitrate of ammonia.

This salt is easily made by the following

Experiment.—Neutralize any convenient amount of ammonia water with nitric acid, and then add to the solution ammonia water until it is slightly in excess. Filter the liquid, and evaporate it carefully to dryness.

$$\underset{\text{Nitric acid.}}{HNO_3} + \underset{\text{Ammonium hydroxide}}{NH_4OH} = \underset{\text{Ammonium nitrate.}}{NH_4NO_3} + \underset{\text{Water.}}{H_2O}.$$

Nitrate of ammonium is a white salt of specific gravity 1.7. It is salty to the taste, soluble in half its weight of cold water, and more soluble in hot water; in either case

solution is accompanied with considerable decrease of temperature. When heated gently it decomposes into nitrous oxide and water, thus:—

$$\underset{\text{Ammonium nitrate.}}{NH_4NO_3} = \underset{\text{Nitrous oxide.}}{N_2O} + \underset{\text{Water.}}{2H_2O}.$$

If suddenly heated it is resolved into nitrogen, water, and nitric oxide (NO). Large amounts of nitrate of ammonium are consumed by dentists for the production of nitrous oxide, or laughing gas, as it is usually called.

Ammonium Oxalate—Formula, $(NH_4)_2C_2O_4$. Molecular weight. 123.8.

Synonyms.—Oxalate of ammonium. Oxalate of ammonia.

Experiment.—Add oxalic acid to a convenient quantity of ammonia water until the ammonia is saturated, then add ammonia water to the solution until it is in slight excess. Evaporate this liquid until a pellicle forms, then place in a cool situation to crystallize. Lastly, drain the crystals in a funnel and dry them on porous paper by exposure to the atmosphere.

Oxalate of ammonium is of value as a reagent, but is not used in medicine. It is a very delicate test for calcium salts in solution, forming with such the very insoluble oxalate of calcium.

Ammonium Phosphate — Formula, $(NH_4)_2HPO_4$. Molecular weight, 131.82.

Synonyms.—Ammonii phosphas. Phosphate of ammonium. Phosphate of ammonia. Hydrogen di-ammonium phosphate.

This salt is occasionally found native, especially in guano. It may easily be made by adding ammonia water to solution of ortho-phosphoric acid until the ammonia is in slight excess, then evaporating the solution and crystallizing. This salt is little used in this country. It is colorless, evolves ammonia when heated with caustic soda or caustic potash, is insoluble in alcohol, soluble in four parts of cold water, and is more soluble in hot water.

Ammonium Succinate.—Formula, $(NH_4)_2C_4H_4O_4$. Molecular weight, 151.74.

Synonyms.—Ammonii succinas. Succinate of ammonium. Succinate of ammonia.

Succinate of ammonium may be made by saturating ammonia water with succinic acid, thus:—

$$\underset{\text{Ammonia.}}{2NH_4OH} + \underset{\text{Succinic acid.}}{H_2C_4H_4O_4} = \underset{\text{Ammonium succinate.}}{(NH_4)_2C_4H_4O_4} + \underset{\text{Water.}}{2H_2O}.$$

Succinate of ammonium is sometimes employed in medicine. It should be white, odorless, and entirely soluble in water. As found on the market, it is often brown and possesses strongly the odor of oil of amber. It precipitates the ferric salts of iron as brownish-red succinate of iron, and is sometimes used to estimate iron in solution.

The German Pharmacopœia recognizes *liquor ammonii succinici*, made as follows:—

Take of succinic acid, powdered, one part, dissolve it in distilled water eight parts, and add of pyrocarbonate of ammonium one part, or sufficient to effect neutralization. Set the liquid aside for twenty-four hours, then filter.

Ammonium Sulphate—Formula, $(NH_4)_2SO_4$. Molecular weight, 131.84.

Synonyms.—Ammonii sulphas. Sulphate of ammonium. Sulphate of ammonia. Normal ammonium sulphate.

This salt is found native in some volcanic districts, especially in Tuscany. The market is at present supplied from the gas works, where it is made by passing ammonia gas into dilute sulphuric acid, and crystallizing the salt after evaporation of the solution. It is at first very impure, and requires several crystallizations to remove the odor of the tar and other products of the destructive distillation of coal.

Sulphate of ammonium forms transparent, odorless crystals of the rhombic system; specific gravity, 1.77 They have a saline taste, dissolve in about two parts of cold water, and in a little more than their weight of boiling water, and are insoluble in alcohol. Sulphate of ammonium is not used in medicine, but is employed, either di-

rectly or indirectly, for the production, in this country, of all other salts of ammonium. It is extensively used as a fertilizer.

Ammonium Sulphide.—Formula, NH_4HS. Molecular weight, 50.99

Synonyms. — Sulphide of ammonium. [Sulphide of ammonia. Sulphuret of ammonia.] Sulph-hydrate of ammonium. Ammonium hydrosulphide.

This is made by saturating ammonia water with sulphide of hydrogen. The process is finished when the solution ceases to cause a white precipitate with solution of sulphate of magnesium (Epsom salt). The reaction is represented as follows:—

$$\underset{\text{Ammonium hydroxide}}{NH_4OH} + \underset{\text{Hydrogen Sulphide.}}{H_2S} = \underset{\text{Ammonium Sulphide.}}{NH_4HS} + \underset{\text{Water.}}{H_2O}.$$

Sulphide of ammonium is at first almost colorless. By age it becomes yellow from oxidation, water and higher sulphides resulting. It has a strong sulphuretted odor, and is an exceedingly valuable reagent, precipitating most of the metals from solution. It should be preserved in small, well closed vials, protected from the light.

Other Sulphides.—There are several other sulphides of ammonium, of little general interest. They may be named as follows:—

Ammonium monosulphide, NH_4S.
Ammonium tetrasulphide, $(NH_4)_2S_4$.
Ammonium pentasulphide, $(NH_4)_2)S_5$.
Ammonium heptasulphide, $(NH_4)_2S_7$.

Ammonium Tartrate. — Formula, $NH_4HC_4H_4O_6$. Molecular weight, 166.65.

Synonyms.—Ammonii tartras. Tartrate of ammonium. Tartrate of ammonia. Acid tartrate of ammonium. Hydrogen tartrate of ammonium.

Tartrate of ammonium is made by adding to ammonia water a solution of tartaric acid until neutralized. Considerable heat will be evolved, and if the solutions are concentrated, crystals will be deposited when the mixture

cools. The reaction is explained by the following equation :—

$$NH_4OH + H_2C_4H_4O_6 = HNH_4C_4H_4O_6 + H_2O.$$

Ammonia water. Tartaric acid. Ammonium tartrate. Water.

This salt may be obtained in the form of a white crystalline powder by evaporating the solution to dryness over a water bath. It is seldom employed in medicine.

Ammonium Valerianate. — Formula, $NH_4C_5H_9O_2$. Molecular weight, 118.78.

Synonyms.—Ammonii valerianas. Valerianate of ammonium. Valerianate of ammonia.

This salt is made by passing ammonia gas into valerianic acid until neutralization is effected. The result after solidification is recrystallized from a concentrated and hot alcoholic solution.

It is a white salt of a greasy or camphor-like appearance, has the disagreeable odor of valerianic acid and a sharp taste. It is dissipated by heat, and without residue. Mineral acids decompose it, valerianic acid separating. Alkaline solutions also decompose it, ammonia being evolved. Valerianate of ammonium is very soluble in water and in alcohol. Its aqueous solution has a strong disagreeable odor, even if the valerianate of ammonium is in very small amount. Upon the contrary, the *alcoholic* solvent completely masks the odor of the valerianate of ammonium, even if the proportion be very great.

Tests for Ammonia.—If the solution of any salt of ammonium be heated with caustic potash, the odor of ammonia is at once evolved.

Experiment.—Mix a little sulphate or chloride of ammonium in a test-tube with a little water, add a small piece of caustic potash and boil. Ammonia escapes, and can be recognized by its odor or by changing to blue a strip of moistened red litmus paper, which is suspended in the mouth of the tube.

2d. If a small amount of "Nessler's test" be added to a solution of ammonia, or a salt of ammonium, a brownish tinge or a brown precipitate results. This test is so delicate as to detect the most minute trace of ammonia.

To Prepare Nessler's Test.— Dissolve 775 grains of iodide of potassium in a small quantity of hot distilled water; continue the heat and carefully add to the solution, with constant stirring, strong hot solution of corrosive sublimate ($HgCl_2$) until the precipitate (mercuric iodide) forming ceases to dissolve. Then filter and mix with the filtrate a concentrated solution of 3100 grains of caustic potash. Add to the above mixture distilled water in sufficient amount to bring it to the bulk of two pints. Then mix with it 1¼ fluid drachms of saturated solution of corrosive sublimate, and after the precipitate subsides, decant the clear solution. This "Nessler's test" is the most delicate known reagent for ammonia. It should be kept in a number of small, securely sealed bottles, as it decomposes by exposure.

METALS OF THE ALKALINE EARTHS.

CALCIUM, BARIUM, STRONTIUM.

In early days the name "earth" was applied to such non-metallic substances as refused to dissolve in water, and remained unchanged upon the application of heat.. Thus the various kinds of clay, which consist of more or less impure silicates of aluminum, as well as the oxides of calcium, barium, etc., were originally associated as *earths.* It was found subsequently that a few, such as lime and magnesia† (oxide of calcium and oxide of magnesium), possessed alkaline properties and seemed to be related to the alkalies proper. They were then called *alkaline* earths, and were considered true elements until Davy demonstrated that they were oxides of various elements.

† Magnesium is considered in the group of elements, embracing beryllium, zinc and cadmium.

CALCIUM.

Symbol, Ca. Atomic weight, 39.9.

This is the most important element of the group, and compounds of calcium are employed very extensively in both in the arts and in medicine, but the *element* is of no practical value. The ancients made lime, and, as we have said, it was long called an earth, and associated under that name with other bodies.

Calcium is never found in a metallic state, but the carbonate under various forms, such as limestone, marble, chalk, coral, etc., occurs naturally in incalculable amounts, and as magnesian limestone (carbonate of calcium and magnesium) it forms the bulk of many mountains.

Calcium salts are necessary to animal life, the bones and teeth of animals consisting mainly of phosphate of calcium. The salts of calcium are always found in plants, and the shells of mollusca and of eggs are very pure specimens of calcium carbonate.

Metallic calcium was obtained in the form of an amalgam by the electrolysis of chloride of calcium in presence of mercury. It was obtained in appreciable amounts by Matthiessen (1856) by heating a mixture of calcium chloride, strontium chloride, and ammonium chloride in a platinum crucible, and decomposing the melted mass with a galvanic battery.

Calcium is yellow, malleable, of specific gravity 1.5778, and tenacious and harder than lead, according to both Bunsen and Matthiessen. But Frey ascribes to it the color of aluminum, and very brittle properties. It remains unaltered in dry air for a considerable period, but in moist air it quickly oxidizes. It burns in the atmosphere when heated to redness, with a yellow light, and decomposes water rapidly if thrown upon it, evolving hydrogen. It is of no importance, excepting to the investigating chemist.

Calcium Monoxide.—Formula, CaO. Molecular weight, 55.86.

Synonyms.—Lime. Caustic lime. Quick lime.

This is made by burning lime-stone (carbonate of calcium) by which means carbon dioxide is driven off, thus:—

$$\underset{\text{Calcium carbonate.}}{CaCO_3} = \underset{\text{Calcium monoxide.}}{CaO} + \underset{\text{Carbon dioxide.}}{CO_2}.$$

It is essential in making lime that a good draught of air pass through the heated stones for the purpose of carrying off the carbon dioxide as fast as it is formed, for if limestone be surrounded by this gas it refuses to part with the carbon dioxide, even at a very high heat.

Calcium monoxide (lime) is white when perfectly pure, but as found in commerce it is of an ash color from impurities pre-existing in the ordinary stone. Its specific gravity is 3.08. When exposed to the air it absorbs both water and carbon dioxide, and crumbles to a powder known as "air-slaked lime." When heated in the flame of an oxy-hydrogen blow-pipe it emits an intense light.† Lime is used to dry gases and to abstract water from alcohol when this is being concentrated to a strength above that yielded by ordinary distillation. It unites with water to form

Calcium Hydroxide, $Ca(OH)_2$, known as slaked lime, or hydrate of lime. If lime is covered with water, or water added to the extent of one-third the weight of the lime, much heat is evolved, and it has been recorded that fires have resulted from this reaction. Slaked lime is pure white; one hundred parts of water at 15.6° C. (60° F.) dissolve 0.1368 parts, and at 100° C. (212° F.) only dissolve 0.0752 parts. The solution in water is known as limewater, and is alkaline both in taste and reaction, it absorbs carbonic acid gas (carbon dioxide) upon exposure to the atmosphere, calcium carbonate precipitating. For this reason limewater should be kept in close vessels. When ordinary slaked lime is rubbed into a thin paste

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† The flame of an oxy-hydrogen blow-pipe cast upon a point of lime forms the *oxy-hydrogen calcium* light. When the stream of oxygen gas is passed through the flame of an alcohol lamp, and then upon a point of lime an inferior light is produced, known as the *oxy-calcium* light.

with water it forms *milk of lime,* which is used in many chemical operations, and for whitening ceilings, etc., under the name of "whitewash."

Chloride of Lime.

Synonyms.—Chlorinated lime. Calcium hypochlorite. Bleaching powder. Calx chlorinata.

Chloride of lime has been known for many years, and several opinions have been advanced concerning its composition. Balard considered it a mixture of calcium hypochlorite, $Ca(ClO)_2$, and calcium chloride, $CaCl_2$, and many at the present day accept that view. Odling advances the idea that it is a double salt or a chlorinated chloride, CaClOCl. Stahlschmidt explains the formation of chlorinated lime upon the assumption that three molecules of calcium hydroxide are decomposed by chlorine, resulting in the formation of water, chloride of calcium, and a compound of the composition $CaHClO_2$, thus:—

$$3Ca(OH)_2 + 2Cl_2 = 2CaHClO_2 + CaCl_2 + 2H_2O.$$

Chloride of lime is made by passing chlorine gas over dry slaked lime, until it ceases to be absorbed. Commercial chlorinated lime contains from 20 to 35 per cent. of chlorine, although it may be produced so as to contain 40 per cent. or even, it is said, 43 per cent. of chlorine. It is white, moistens upon exposure to the air and decomposes, absorbing, in addition to water, carbon dioxide. It has a strong chlorine odor under such circumstances. When freshly prepared it has an odor of hypochlorous acid. Even if preserved in closed vessels it decomposes, and it has been recorded that explosions have occurred, but the reaction is unexplained. According to Kingzett a strong aqueous solution of chloride of lime deposits crystals of calcium hypochlorite, $Ca(ClO)_2 + 4H_2O$, when exposed over sulphuric acid in the exhausted receiver of an air pump. Chloride of lime is used as a deodorizer and disinfectant, to bleach cotton and other goods, to make chloroform and the officinal solution of chlorinated soda.

Calcium Dioxide, CaO_2, is formed by passing carbon dioxide into limewater. It precipitates in combination with eight molecules of water (CaO_2+8H_2O), which give off water of crystallization at a temperature of 130° C. (266° F.) This compound is of no practical value.

Calcium Bromide.—Formula, $CaBr_2$. Molecular weight, 199.40.

Synonyms—Calcii bromidum. Bromide of calcium.

This salt is most easily prepared by saturating hydrobromic acid with prepared or precipitated chalk, and after filtration evaporating the solution to dryness. Bromide of calcium is a white salt, soluble in alcohol and in water and of a saline bitterish taste. It is scarcely liable to adulteration, being so little used as to hardly make it an object.

Calcium Carbonate. — Formula, $CaCO_3$. Molecular weight, 99.75.

Synonyms.—Calcii carbonas. Carbonate of calcium. Carbonate of lime.

This substance is found in unlimited amounts over the entire globe. It forms vast layers in the ocean, derived mainly from the shells of mollusca, and beneath the surface soil of inland countries it is not unusual to find the layers of limestone perfect conglomerates of both broken and well preserved petrefactions of extinct sea shell-fish. This is forcibly illustrated by the region of country in the author's neighborhood, for each layer of limestone of the hills and valleys near Cincinnati is a mass of petrefactions. The immense number of these layers of stone, one above the other, speaks of the great time that must have passed before such quantities of carbonate of lime could have formed from this origin. Coral reefs are carbonate of calcium; many mountain ranges and the chalk formations, the marble cliffs and other forms of limestone, are *carbonate of calcium.* Calcium carbonate is dimorphous, but it is unnecessary to consume space with a history of its crystalline modifications. When a soluble calcium salt is precipitated by the carbonate of an alkali

metal, the precipitate formed is carbonate of calcium and thus we have the officinal—

Calcii Carbonas Præcipitata, or precipitated carbonate of calcium, made by adding a hot solution of carbonate of sodium to a hot solution of chloride of calcium, washing the precipitate and drying it. The reaction is expressed as follows:—

$$\underset{\text{Calcium chloride.}}{CaCl_2} + \underset{\text{Sodium carbonate.}}{Na_2CO_3} = \underset{\text{Calcium carbonate.}}{CaCO_3} + \underset{\text{Sodium chloride.}}{2NaCl.}$$

Precipitated chalk is a very white semi-crystalline powder and should not be confused with—

Creta Præparata, or prepared chalk, which is made by rubbing chalk into an impalpable powder, mixing it with water and then, after allowing the coarse particles to settle, decanting the liquid, permitting it to settle and drying this last precipitate. Prepared chalk is usually of slight ash color. It is customary to dry it in conical drops. This form of chalk is heavier, bulk for bulk, than precipitated chalk.

Properties.—Carbonate of calcium is very slightly soluble in either cold or boiling water, but is less soluble when ammonia or carbonate of ammonium is present. It is soluble to a considerable extent in carbonic acid water, from which solution it is deposited by exposure to the air or by boiling. This fact accounts for the deposit of carbonate of calcium inside of steam boilers where "hard" water is used, and for the natural crystals of carbonate of calcium, although often in the latter case immense periods of time may have been necessary for their formation. All of the forms of carbonate of calcium dissolve with effervescence in dilute mineral acids, carbon dioxide escaping, although sulphuric acid speedily forms a layer of *insoluble* sulphate of calcium over the surface of the lumps which impedes the reaction.† Aside from its use in

† We can say that the stick chalk, for blackboard use, which we have examined, instead of being carbonate of calcium, was sulphate of calcium. This is of no importance for the use intended, but sulphate of calcium has been also substituted for precipitated carbonate of calcium, and for prepared chalk.

medicine, carbonate of calcium is of incalculable value in the arts, as for building purposes, iron production, etc.

Calcium Chloride.—Formula, $CaCl_2$. Molecular weight, 110.64.

When ammonia water is made by decomposition of chloride of ammonium and calcium hydroxide, this compound results, thus:—

$$\underset{\text{Ammonium chloride.}}{2NH_4Cl} + \underset{\text{Calcium hydroxide.}}{Ca(OH)_2} = \underset{\text{Calcium chloride.}}{CaCl_2} + \underset{\text{Ammonium hydroxide.}}{2NH_4OH.}$$

Chloride of calcium is found in the water of some springs and in sea-water. It is also produced on a large scale as an incidental product in the manufacture of potassium chlorate. It may be made by dissolving chalk or fragments of marble in hydrochloric acid, passing chlorine through the solution in order to peroxidize the iron and manganese, and then adding milk of lime in slight excess, filtering and evaporating the filtrate to dryness over a fire.

Chloride of calcium may be obtained in the form of crystals, but as usually found on the market, is in amorphous lumps. It has a strong affinity for moisture, and is employed by chemists in drying gases. For this purpose the commercial article is unsuited, owing to the large amount of water present. This may be properly prepared by placing it in a porcelain or even an iron evaporating dish, and heating over a fire until it has melted and is dry again, and then while hot transferring it to well closed glass-stoppered vials. Chloride of calcium is used in the preparation of artificial mineral waters.

Calcium Citrate.—Formula, $Ca_3 2C_6H_5O_7$. Molecular weight, 496.78.

Citrate of calcium is produced in immense quantities, but is of interest from the fact that it gives rise to *citric acid*, rather than that it is of value as a compound. By referring to our article on citric acid the manner of preparing citrate of calcium may be noted, thus rendering its reproduction unnecessary in this place.

Citrate of calcium is one of the few salts that dissolves to a greater extent in cold than in hot water. Advantage is taken of this fact to identify citric acid, for if lime-water be added in excess to a solution of citric acid, a precipitate will be produced when the solution is boiled, which will redissolve upon cooling. Citrate of calcium is made in countries which produce limes and lemons in abundance, and from thence is exported for the purpose of preparing citric acid. Gum Arabic is mostly a compound of calcium, and an acid known as *gummic acid.* (See gums.)

Calcium Hypophosphite. — Formula, $Ca2H_2PO_2$. Molecular weight, 169.66.

Synonyms— Calcii hypophosphis. Hypophosphite of calcium. Hypophosphite of lime.

Hypophosphite of calcium may be made by the following—

Experiment.—Add carbonate of calcium to hypophosphorous acid until the acid is saturated and effervescence ceases. Filter the solution and evaporate the filtrate to dryness over a water bath. The reaction is expressed as follows:—

$$\underset{\text{Hypophosphorous acid.}}{2H.H_2PO_2} + \underset{\text{Calcium carbonate.}}{CaCO_3} = \underset{\text{Calcium hypophosphite.}}{Ca2(PO_2H_2)} + \underset{\text{Carbon dioxide.}}{CO_2} + \underset{\text{Water.}}{H_2O}.$$

In the practical preparation of hypophosphite of calcium a different process is employed. Hypophosphite of calcium is the first of a series of salts of hypophosphorous acid, and by reference to our article on hypophosphorous acid it will be seen that it is used as a basis for preparing that acid, and therefore the acid cannot be used with economy for reproducing the calcium salt. In a practical way calcium hypophosphite is made by acting upon phosphorus with milk of lime. The reaction is complicated, phosphide of hydrogen, hydrogen and phosphoric acids also being formed. In consequence of this the operation must be performed in the open air or under a good flue, and even then the experiment is not pleasant.

Hypophosphite of calcium crystallizes in four sided

prisms. It is soluble in six parts of cold water, and is *less* soluble in hot water, therefore it is not found in market in definite crystals, but either as a granular powder, which is made by evaporating the solution to dryness, or as pearly scales, which are prepared by evaporating the solution and skimming the crust as it forms over the surface of the liquid. Hypophosphite of calcium has a saline bitterish taste. When heated it decomposes, at a certain temperature evolving phosphide of hydrogen, water, and phosphorus, sometimes with explosions. Hypophosphite of calcium is liable to contain caustic lime as an impurity, and we have found it necessary to remove this article from the commercial salt. In preparing hypophosphite of calcium, occasionally a little hypophosphorous acid should be added to the solution while it is evaporating.

Calcium Iodide.—Formula, CaI_2. Molecular weight, 292.96.

Synonyms.—Calcii iodidum. Iodide of calcium. Iodide of lime.

This salt is easily made as follows:—

Experiment.—Place a convenient amount of hydriodic acid in a porcelain capsule, and add an excess of calcium carbonate and stir until the solution is neutral. Filter this and evaporate the filtrate to dryness. Iodide of calcium is a white, deliquescent powder, which decomposes upon exposure to the atmosphere. It is very soluble in water and dissolves in alcohol. Iodide of calcium must not be confused with *iodate* of calcium ($Ca2IO_36H_2O$) which, as yet, has no therapeutical value.

Calcium Oxalate.—Formula, CaC_2O_4. Molecular weight, 127.68.

Synonyms.—Calcii oxalas. Oxalate of calcium. Oxalate of lime.

This salt is not used in medicine, but is of interest to the chemist. When a solution of oxalate of ammonium is added to any neutral solution of a salt of calcium, an immediate precipitate of oxalate of calcium is produced. This precipitate remains upon the addition of an excess

of acetic acid, but either nitric acid or hydrochloric acid will readily dissolve it. Oxalate of calcium is white, tasteless, insoluble in water and alcohol. It is a by-product of the manufacture of hypophosphorous acid, and as thus obtained is in a very finely divided state, and is particularly valued for the purpose of brightening and polishing silverware.

Calcium Phosphate.—Formula, $Ca_3(PO_4)_2$. Molecular weight, 229.5.

Synonyms. — Calcii phosphas præcipitata. Precipitated phosphate of calcium. Phosphate of calcium. Phosphate of lime. Normal calcium ortho-phosphate. Bone-phosphate.

Calcium ortho-phosphate occurs native in several localities, and some minerals consist of nearly pure calcium ortho-phosphate, as coprolites (probably of animal origin) and the mineral osteolite. It constitutes the larger part of the inorganic portion of bones, fully 80 per cent. of bones burned in the air to whiteness consisting of calcium phosphate. It is a constituent of vegetable organisms, being abundant in fruits and seeds. It is a recognized fact that the flour from wheat which has been stripped of its outer coatings, in order to produce a very white flour, is of poorer quality than when the grain is ground entire. This is mainly owing to a deficiency of calcium phosphate which is most abundantly deposited in the outer layers of the grain.

Phosphate of calcium may be obtained as a white powder by precipitating a solution of calcium chloride mixed with ammonia water by means of a solution of sodium phosphate, and washing the precipitate well and then drying it. According to the U. S. P. it is made by dissolving burnt bones (bone-ash) in muriatic acid, diluting and filtering the solution, and then precipitating it by means of ammonia water. The precipitate is to be washed with water until the washings cease to effect a solution of nitrate of silver acidulated with nitric acid, and is then to be dried. As made by this process it is contaminated with calcium fluoride and magnesium phosphate, but the

amount of these impurities is insufficient to interfere with its therapeutic action.

Phosphate of calcium is almost insoluble in water, but it slowly decomposes in cold water, giving rise to an insoluble basic salt and a soluble acid salt, and this change is favored by boiling. Therefore, according to Warrington, the exact solubility of calcium ortho-phosphate can not be determined. It dissolves to an extent in solutions of ammonium salts, and the acids, almost, if not all of them, dissolve it, and usually very freely, even carbonic acid solution may not be excepted.

There are other calcium phosphates, but none of use in medicine. The *mono-hydrogen calcium ortho-phosphate* ($CaHPO_4$) is obtained when a solution of calcium phosphate is mixed with a solution of phosphate of sodium, and is sometimes deposited from urine as concretions or in small crystals.

Tetra-hydrogen calcium phosphate, $H_4Ca(PO_4)_2$, is obtained by acting upon bone-ash with sulphuric acid, and the mixture of this salt with the calcium sulphate also formed, is sold under the name *super-phosphate of lime*, and is very largely used as a fertilizer.

Calcium Sulphate. — Formula, $CaSO_4$. Molecular weight, 135.72.

Synonyms.—Calcii sulphas. Sulphate of calcium. Sulphate of lime.

This compound occurs native in a number of forms, such as gypsum ($CaSO_4+2H_2O$), of which the large crystals, selenite, the satin-spar or fibrous gypsum, and alabaster, are modifications. Gypsum loses its water of crystallization at the temperature of 110° to 120° C. (230° to 248° F.) and forms a white substance which, when powdered, is called *plaster of Paris* well known for its property of combining again with water and becoming hard. Sulphate of calcium is slightly soluble in cold water, but dissolves to a greater extent in water at 35° C. (95° F.) than at higher or lower temperatures. The modification of this salt made by decomposing solution of calcium chloride by means of sulphuric acid is known

as pearl-hardening or analine, and is used to give weight and gloss to paper. Plaster of Paris should be either fresh or preserved in glass stoppered bottles, as it loses its power of hardening after exposure to the air for any length of time, and as a consequence is then unreliable for use in surgery where it is frequently employed to obtain casts of deformities, and occasionally to hold broken limbs in position.

Calcium Sulphite.—Formula, $CaSO_3$. Molecular weight, 119.76.

Synonyms.—Calcii sulphis. Sulphite of calcium. Sulphite of lime.

Sulphite of calcium precipitates as a white powder when a solution of sulphite of sodium is mixed with a solution of chloride of calcium. It is usually made in commerce by passing sulphur dioxide over layers of slaked lime made into a paste with water, and although impure in this instance, answers for the purposes required of it. Sulphite of calcium is white, has a faint sulphurous taste, and absorbs oxygen and changes by exposure to the atmosphere into sulphate of calcium. It dissolves in about 800 parts of cold water, and dissolves very freely in solution of sulphurous acid forming what is known in commerce as *bisulphite of lime.* Sulphite of calcium is not used in medicine, but is employed to preserve wines and cider from fermentation, and we have known the solution of it in sulphurous acid to be used for preserving fruit.

Calcium Monosulphide.—Formula, CaS. Molecular weight, 71.88.

Synonyms.—Calcii monosulphidum. Calcium sulphide. Sulphide of calcium. Sulphuret of lime.

This may be made by heating a mixture of charcoal and sulphate of calcium, or by calcining calcium hydroxide or calcium carbonate with sulphur. In this latter case sulphate of calcium is also formed. Sulphide of calcium is a white powder, insoluble in water, odorless when first prepared, but after exposure to a moist atmosphere it

gives off the odor of sulphide of hydrogen. It dissolves in dilute hydrochloric acid with evolution of sulphide of hydrogen. When exposed to the sunlight it is for some time afterward luminous in the dark, and the old (1768) so-called "Canton's phosphorus" is sulphide of calcium made by calcining oyster shells with sulphur.

Calcium Disulphide. — Formula, CaS_2. Molecular weight, 103.86.

Synonyms.—Disulphide of calcium. Sulphide of calcium. Sulphuret of lime (or calcium).

This may be made by boiling a mixture composed of sulphur two parts, lime one part, and water sixteen parts, until the liquid assumes a reddish yellow color, then filtering. From this by gentle evaporation and subsequent cooling yellow crystals will separate of the composition $CaS_2 3H_2O$. Calcium disulphide is used in medicine in the form of a solution and is never employed in crystals. It is commonly known as solution of sulphuret of calcium and has attained considerable reputation as a remedy for the itch. The most serious objection to it is the offensive odor.

Calcium Tartrate.—Formula, $CaC_4H_4O_6$. Molecular weight, 187.54.

Synonyms.—Calcii tartras. Tartrate of calcium. Tartrate of lime.

This salt is scarcely of use in medicine, but interests the chemist from the fact that by its behavior the distinction can be drawn between citric and tartaric acids. If to a solution of any normal tartrate, chloride of calcium solution be added in slight excess, a white precipitate of tartrate of calcium will at once appear, and upon collecting this precipitate and adding it to a cold solution of caustic potash it dissolves, upon warming the solution it re-precipitates. Citrate of calcium under like conditions will not dissolve in the cold potash solution. It is necessary to guard against the presence of ammonium salts in the foregoing experiment, as citrate of calcium is soluble in such solutions.

Tests for Calcium.—Solution of carbonate of ammonium precipitates calcium salts from solution as *carbonate* of calcium. Solution of oxalate of ammonium precipitates calcium salts from solution as *oxalate* of calcium. This last precipitate when fresh will dissolve readily upon addition of nitric or hydrochloric acid, but it is insoluble in acetic acid. These precipitates are not conclusive evidences of calcium, as barium and strontium salts react in like manner. If barium and calcium salts are mixed, yellow chromate of potassium will precipitate the barium, and from the filtered solution the calcium may be obtained with the foregoing reagents.

If a calcium salt be moistened with hydrochloric acid, and then placed in the hottest part of a Bunsen flame, a red color results. The spectrum of this flame is characteristic and decisive.

STRONTIUM.

Symbol, Sr. Atomic weight, 87.2.

The village of Strontian, in Argyllshire, gave to this element its name, from the fact that it was produced from a mineral found in that locality. The chief sources of strontium are sulphate of strontium, or celestine, $SrSO_4$, and carbonate of strontium, $SrCO_3$. It occurs in small quantities in some other minerals, and the sulphate and chloride of strontium are found in some salt springs and other mineral waters, and in the ashes of the sea-weed *Fucus vesiculosus.*

The element was first obtained by Davy in 1808, by electrolysis of one of its salts. It is rare and of interest at the present day, only to experimental chemists.

Properties.—Strontium is yellow, malleable and harder than either cadmium or lead. Its specific gravity is 2.5. It melts at redness, oxidizes quickly upon exposure to the atmosphere, decomposes water, and burns if heated. None of the salts of strontium are used in medicine. Strontium salts color the flame of an alcohol lamp, or of the blowpipe to a magnificent crimson. Nitrate of strontium is

employed to make red signal lights and the crimson or red fireworks.

Tests for Strontium.—The spectroscope will detect minute amounts of this metal. Carbonate of ammonium precipitates strontium from solution when its salt is mixed with salts of the alkali metals. If calcium salts be present at the same time, a precipitate of calcium carbonate also forms. To detect strontium in such cases, dissolve the precipitate, which forms with ammonium carbonate, in hydrochloric acid, and add to it a solution of sulphate of calcium. If strontium be present, after some time, a turbidity will follow. Nitrate of strontium refuses to dissolve in absolute alcohol. Nitrate of calcium is soluble.

BARIUM.

Symbol, Ba. Atomic weight, 136.8.

The natural source of barium is sulphate of barium ($BaSO_4$) or heavy spar, and carbonate of barium ($BaCO_3$) or witherite, although many ores of manganese, certain silicates, felspathic rocks, mineral waters and sea water contain it in greater or less amounts. Davy first obtained the element, although its preparation was very difficult. The metal oxidizes quickly and burns when heated in the air. According to Frey it requires a greater heat to melt barium than is necessary to fuse cast-iron. Very few of the salts of barium are used in the arts or in medicine.

Barium Monoxide.—Formula, BaO. Molecular weight, 152.76.

Synonyms.—Monoxide of barium. Baryta.

This is made by heating nitrate of barium in a crucible until red fumes cease to be evolved. Baryta has an intense affinity for water, exceeding that of lime. If moistened with water it combines with such violence, and liberates so much heat, as to bring the lump to redness, forming—

Barium Hydroxide, $Ba(OH)_2$, or common caustic baryta. This is prepared, however, in practice by heat-

ing barium sulphide, passing over it a current of moist carbon dioxide as long as it is absorbed and afterward subjecting the carbonate formed, while it is hot, to a current of superheated steam. The carbon dioxide and water react upon the barium sulphide to form barium carbonate and hydrogen sulphide, the latter escaping, thus:—

$$\underset{\text{Barium sulphide.}}{BaS} + \underset{\text{Carbon dioxide.}}{CO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Barium carbonate.}}{BaCO_3} + \underset{\text{Hydrogen sulphide.}}{H_2S}.$$

Afterward, the superheated steam decomposes this carbonate with evolution of carbon dioxide and formation of barium hydroxide, thus:—

$$\underset{\text{Barium carbonate.}}{BaCO_3} + \underset{\text{Water.}}{H_2O} = \underset{\text{Barium hydroxide.}}{Ba(HO)_2} + \underset{\text{Carbon dioxide.}}{CO_2}.$$

Barium hydroxide is a white powder which melts at a low red heat and has the specific gravity 4.495. It unites with water to form a crystalline hydrate $BaO_2H_2+8H_2O$, which is soluble in water, and which gives off seven molecules of water upon exposure to a dry atmosphere, changing into a white powder. When hydroxide of barium is dissolved in water it forms the common *baryta-water*, used in analytical work to absorb carbon dioxide, for hydroxide of barium has an intense affinity for that gas. Its solution in water is strongly alkaline. Barium hydroxide is not used in medicine, but is employed in sugar refining, and is very valuable to the experimental chemist.

Barium Chloride.—Formula, $BaCl_2$. Molecular weight, 207.54.

Synonyms.—Barii chloridum. Chloride of barium. Muriate of baryta.

This salt is made by dissolving witherite (native carbonate of barium) in dilute hydrochloric acid, and afterward purifying the chloride of barium from calcium, lead, iron, and manganese salts which are present in that mineral. It is not essential to notice the process farther in this work, as the salt is never made on a small scale. It is also made from native sulphate of barium. Barium chloride crystallizes in transparent rhombic plates with

two molecules of water, $BaCl_2 + 2H_2O$. They are permanent, have the specific gravity 3.05, dissolve in about four parts of cold water and in less than twice their weight of boiling water, and are insoluble in alcohol. Solution of barium chloride is precipitated by a soluble sulphate or by sulphuric acid, and the precipitate is insoluble in strong nitric acid. An officinal solution of chloride of barium is made by dissolving one troy ounce of barium chloride in three fluid ounces of water, but it is seldom used.

Barium Nitrate. — Formula, $Ba(NO_3)_2$. Molecular weight, 260 58.

This salt is made in quantities, being consumed for pyrotechnic purposes, as the flame produced when it is mixed with combustible bodies is of a beautiful green. It is made from either the sulphide of barium or the natural carbonate, crystallizes in transparent crystals, has an acid taste, is soluble in water, insoluble in alcohol and does not enter into the composition of a single remedial agent.

Barium Sulphate.—Formula, $BaSO_4$. Molecular weight 232.62.

This is the common source of the barium compounds, and is known under the name *heavy-spar*. It is used in immense amounts for the purpose of adulterating white lead, and is called baryta or barytes in commerce. For this purpose, however, artificial sulphate of barium is preferred, as it has more "*body*" than the natural crystals. That is, it is opaque, whilst the powdered heavy-spar is partly transparent. It is produced artificially by precipitating a solution of chloride of barium with dilute sulphuric acid. Sulphate of barium is, practically, insoluble in water.

All of the soluble salts of barium are *powerful poisons*. The proper antidotes are solutions of the sulphates of sodium (Glauber's salt), or magnesium (Epsom salt), or aluminum.

Tests for Barium.—Volatile salts of barium impart a greenish color to an alcohol flame. If an insoluble ba-

rium salt is moistened with sulphuric or hydrochloric acid and held on a loop of platinum wire in a Bunsen gas flame, the part of the flame just over it will become of a yellowish-green. If this flame be viewed through green glass, it appears bluish-green. When soluble barium salts are mixed with soluble salts of calcium and strontium; to separate the barium, add solution of calcium sulphate until it ceases to cause a precipitate. The barium is precipitated while the other elements remain in solution. If the sulphates of calcium, strontium and barium are mixed, treat the mixture for twenty-four hours with a concentrated solution of ammonium carbonate. Then digest the residue with hydrochloric acid; strontium and calcium are removed while sulphate of barium remains.

MAGNESIUM GROUP OF METALS.

BERYLLIUM, CADMIUM, MAGNESIUM, ZINC.

These elements burn when heated in the air, and are volatile. Their sulphates are soluble, and each metal forms one oxide. Excepting beryllium, certain of their salts are more or less used in medicine.

BERYLLIUM (*Glucinum*).

Symbol, Be. Atomic weight, 9.0.

Beryllium oxide was discovered in 1798 by Vauquelin, who demonstrated that it was present in *beryl*, a mineral, which before that time was considered a compound of lime or alumina. The name *glucinum* was given to it afterward in consequence of the sweet taste of its salts, but was not accepted, and the favored name at present is *beryllium*. Wœhler first obtained the element (1828), as a dark gray powder. Debray afterward obtained it in mass.

Beryllium is of silver brightness, white, of specific gravity 2.1, and melts at a less temperature than silver. None of its salts are used in medicine.

CADMIUM.

Symbol, Cd. Atomic weight, 111.6.

Stromeyer discovered cadmium in 1817, obtaining it from carbonate of zinc ore, and almost simultaneously Hermann identified it in the same compound of different origin. It is found in various zinc ores, and is separated by suitable means. It is of interest mainly to mineralogists.

Cadmium is white and presents a fibrous fracture. It can be rolled into foil, drawn into wire, and is harder than tin, giving out when bent, a crackling sound, as that metal does in the act of bending. Its specific gravity is 8.546 when cooled from the molten state, but 8.667 after hammering.

Cadmium Iodide.—Formula, CdI_2. Molecular weight, 364.66.

Synonyms.—Cadmii iodidum. Iodide of cadmium.

This compound may be easily made by digesting iodine, water, and an excess of metallic cadmium filings or shavings in a warm situation until the solution is colorless and the iodine has disappeared. Then evaporate and filter it. If the water be in limited amount, crystals of iodide of cadmium will appear before the reaction ends, and when the solution becomes colorless, the sides of the vessel and the excess of cadmium will be covered with beautiful satin-like plates, resembling silvery fish-scales.

Iodide of cadmium may also be made by dissolving metallic cadmium in hydriodic acid, evaporating the solution and crystallizing it

Iodide of cadmium is of a satin-silver color, permament, soluble in alcohol and in water, and of specific gravity 4.576. It is extensively used in photography, and occasionally in medicine, usually in this latter instance, in form of ointment.

Cadmium Sulphate. — Formula, $CdSO_4$. Molecular weight, 207.42.

Synonyms.—Cadmii sulphas. Sulphate of cadmium.

This salt is easily made by acting upon metallic cadmium with a somewhat dilute mixture of sulphuric and

nitric acids, and evaporating to crystallization. It is also made according to the U. S. P. by forming a solution of nitrate of cadmium and precipitating this with carbonate of sodium, and dissolving this precipitate with sulphuric acid and evaporating the solution to crystallization. Sulphate of cadmium is seldom employed in medicine. It is very astringent, and like the other salts of cadmium, acts as an emetic, even in small doses. Iodide and sulphate of cadmium are the only salts of this metal employed in medicine.

Tests for Cadmium.—The alcohol flame is not colored by any salt of cadmium. Pass sulphide of hydrogen into the solution of a cadmium salt, yellow sulphide of cadmium results. If other yellow sulphides are also formed, such as sulphides of antimony, tin and arsenic, add solution of caustic alkalies. Cadmium sulphide refuses to dissolve, the others form solutions. If copper be present, decant the solutions, and mix the precipitate with solution of potassium cyanide. Sulphide of copper dissolves, sulphide of cadmium remains as a yellow magma. For thorough instructions for detection of cadmium in mixtures, see some work devoted to analysis.

MAGNESIUM.

Symbol, Mg. Atomic weight, 23.94.

Magnesium was obtained in an impure state by Davy. Bussy prepared it pure. The compounds of magnesium are very abundantly distributed, both in mineral formations and in saline waters. The element is now manufactured on a large scale in this country and in Europe, and is consumed in considerable amounts by analysts, and for producing the magnesium light. Magnesium is silver-white, permanent in dry air, but oxidizes upon the surface in moist air. It is very light, having a specific gravity of but 1.75, melts at redness, if excluded from the atmosphere, and burns when heated in the flame of an alcohol lamp or a candle, the light being intense and brilliant. Upon this account it has been recommended for lighting purposes, but difficulties have been experienced, such as

in regulating the flow of the wire and in guarding against the fine dust formed during the combustion (oxide of magnesium). It has been estimated that 2½ ounces of magnesium wire will evolve as much light as 20 pounds of stearin candles. This light is very valuable for photographing in caverns, etc., as it is rich in the chemically active rays. The compounds of magnesium are extensively employed in medicine and have been since a very early day.

Magnesium Oxide.—Formula, MgO. Molecular weight, 39.9.

Synonyms.—Magnesia calcinata. Calcined magnesia. Magnesia. Magnesia usta.

This compound is the result of the combustion of magnesium in the atmosphere. It may be made by calcining a magnesium salt, the acid of which is volatile. According to the Pharmacopœia it is made by calcining carbonate of magnesium until effervescence ceases, when a portion of it is dropped into dilute acids. In this reaction carbon dioxide escapes and oxide of magnesium remains, thus:—

$$\underset{\text{Magnesium carbonate.}}{MgCO_3} = \underset{\text{Magnesium oxide.}}{MgO} + \underset{\text{Carbon dioxide.}}{CO_2}.$$

Magnesia is a white powder, very fine and light, when made by the officinal process, almost insoluble in water, tasteless, and unites with acids to form salts, the most of which are soluble. If *heavy* carbonate of magnesium be employed in its preparation the resulting oxide will be of identical composition but much heavier, and thus in commerce we have the two forms, such being distinguished by the names, heavy calcined magnesia, and light calcined magnesia. The tests for purity of magnesia are the same as for magnesium carbonate.

Magnesium Hydroxide, $Mg(OH)_2$, is a compound made by precipitating a solution of a magnesium salt with caustic potash or caustic soda. It is a white powder, but is not recognized as a therapeutic agent.

Magnesium Carbonate.—Formula, $MgCO_3$. Molecular weight, 83.79.

Synonyms.—Magnesii carbonas. Carbonate of magnesium. Carbonate of magnesia.

If a solution of sulphate or of chloride of magnesium be precipitated with a solution of sodium carbonate, and the mixture boiled, and the precipitate afterward washed with cold water and dried, the article known as *light* carbonate of magnesium (magnesia alba levis), is obtained. If the mixed solutions be evaporated to dryness, and then the sulphate or chloride of sodium be dissolved from the mass by means of hot water, the precipitate when again dried will be dense, and is known as *heavy* carbonate of magnesium (magnesia alba ponderosa). Both are identical in chemical composition, in reality being *basic* carbonate of magnesium united with from two to three molecules of water and represented by the formula—

$3MgCO_3Mg(HO)_2.2H_2O$; or, $3MgCO_3.Mg(HO)_2.3H_2O$.

According to R. and S., as found in the shops, it is usually represented by the first formula. Patterson's carbonate of magnesia is familiar in this country, and is made by a patented process which consists in igniting dolomite (a mixture of carbonates of magnesium and calcium) and extracting the result with water saturated with carbonic acid gas. This dissolves the carbonate of magnesium, but scarcely affects the carbonate of calcium, and from the solution the carbonate of magnesium is precipitated by a current of steam. Carbon dioxide passed into water in which carbonate of magnesium is suspended, effects its solution, and the *liquor magnesiæ carbonatis* of the British Pharm. (fluid magnesia) is such a solution representing about 13 grains of carbonate of magnesium to each fluid ounce. Carbonate of magnesium is soluble in solutions of ammonium salts, but is scarcely affected by pure water. It dissolves in diluted acids with effervescence, and the solution made with dilute nitric acid and filtered should not yield a precipitate with barium chloride (absence of sulphates), or with silver nitrate (absence of chlorides and bromides), or with ammonium

oxalate (absence of calcium). It is produced in immense amounts in England directly, and as incidental products. By either process it is difficult to obtain a chemically pure carbonate, and therefore, it can not be expected that a commercial article will answer all the tests mentioned.

Magnesium Chloride,—Formula, $MgCl_2$. Molecular weight, 94.68.

Synonyms.—Magnesii chloridum. Chloride of magnesium.

This salt is obtained in quantities at Stassfurt, where it is found in the saline beds in combination with calcium, and is known as *tachhydrite*, ($2MgCl_2,CaCl_2+12H_2O$). It is also a component of sea-water, many salt springs, and in salt deposits. It may be made by way of experiment by dissolving either carbonate or calcined magnesia in hydrochloric acid previously diluted with its bulk of water, filtering the solution and then evaporating it to crystallization.

Chloride of magnesium dissolves in less than its weight of cold water, and in somewhat more than one-fourth its weight of water at the boiling point of the saturated solution. It crystallizes in crystals of the formula $MgCl_2.6H_2O$, which are deliquescent. It decomposes upon heating, hydrochloric acid and water being evolved, and magnesia remaining. It forms crystallizable double chlorides with many other chlorides.

Magnesium Citrate.—Formula (neutral), $Mg_3 2C_6H_5O_7$. Molecular weight, 448.9. Crystallized—$Mg_3 2C_6H_5O_7.14H_2O$. Molecular weight, 700.34. Acid salt, or hydrogen magnesium citrate, $MgHC_6H_5O_7$. Molecular weight, 213.48.

Synonyms.—Magnesii citras. Citrate of magnesium. Citrate of magnesia.

The hydrogen magnesium citrate is made by the following experiment:—

Magnesium carbonate is mixed with water in a capacious porcelain capsule, and the mixture is warmed to the temperature of 82° C. (179.6° F.) and into it is then stir-

red solution of citric acid until the carbonate of magnesium is dissolved, and effervescence ceases. Add a little carbonate of magnesium to the solution, in order that all free citric acid may be neutralized, after which the solution is to be filtered and evaporated to dryness in a current of air, at the temperature of about 82° C. (179.6° F.) The normal salt is much less soluble than the hydrogen magnesium citrate, and is precipitated from solutions of the foregoing upon standing. Neither salt is used in the dry form in this country, but a preparation known as—

Solution of Citrate of Magnesia, is a favored pharmaceutical, and in much demand. According to the present officinal process, bicarbonate of potassium is employed to yield the required carbon dioxide, although in making solution of citrate of magnesium upon a large scale, freshly precipitated carbonate of magnesium is dissolved in carbonic acid water, under pressure of the carbonic acid gas, and this solution is poured upon the proper amount of syrup of citric acid, previously placed in the "magnesia bottles." These are then corked, tied over with strings, and agitated, the result being formation of citrate of magnesium with liberation of carbon dioxide.†

Magnesium Sulphate.—Formula. $MgSO_4$. Molecular weight, 119.76. Crystallized— $MgSO_4 7H_2O$. Molecular weight, 245.48.

Synonyms. Magnesii sulphas. Sulphate of magnesium. Sulphate of magnesia. Epsom salt.

Sulphate of magnesium is the most common, and most extensively used of the salts of this metal. It occurs native in many minerals, in the deposits at Stassfurt, in caves throughout the United States and other countries, in gypsum quarries, and in most mineral springs. Sulphate of magnesium is produced upon a large scale in all civilized countries, and is so cheap in the market, as to

† In practice it is customary to place the solution of carbonate of magnesium in the bottles, and then through a funnel with a long glass exit tube, to pour the syrup of citric acid upon the bottom of the bottle beneath this solution.

render its preparation by pharmacists unnecessary, unless for an experiment. It may be made in an experimental manner, by dissolving either the carbonate of magnesium or the calcined magnesia in dilute sulphuric acid, and afterward evaporating the solution to crystallization.

Common Epsom salt is in the form of small, colorless, needle-like crystals, usually slightly moist, and are represented by the formula, $MgSO_4.7H_2O$. They are made by dissolving crude sulphate of magnesium in hot water, almost to saturation, and then rapidly cooling the solution. If they absorb moisture, chloride of magnesium is present. Owing to the cheapness of this salt it can not be adulterated with any object, and impurities of a serious character are not likely to occur. The salt resembles, somewhat, oxalic acid, and very closely, in general appearance, sulphate of zinc. Through carelessness they could be mistaken.† It dissolves in less than its own weight of boiling water, and in one and one half parts of cold water, but it is insoluble in alcohol. Upon heating the crystals, they part with six molecules of water at a temperature of 150° C. (302° F.) but a temperature of 200° C. (392° F.) is required to drive off the remaining molecule.

Sulphite of Magnesium, $MgSO_3,6H_2O$., may be made by dissolving either the carbonate of magnesium or calcined magnesia in solution of sulphurous acid, and then evaporating the solution to crystallization in a current of cool air by exposure in a shallow evaporating vessel. It is very seldom employed.

Tests for Magnesium.—Add to the solution of a salt of magnesium, a little ammonium chloride and then ammonium carbonate, no precipitate is produced. Add to this solution a solution of phosphate of sodium (Na_2HPO_4); a precipitate ($MgNH_4PO_4$) follows, either at once or after stirring the mixture with a glass rod. If a precipitate

† A blunder of this nature could be detected by testing with solution of calcium chloride for oxalic acid, a white precipitate identifying that compound. Sulphate of zinc will give a precipitate with sulphide of ammonium or with ferrocyanide of iron and sulphate of magnesium will not.

follows the addition of the ammonium carbonate, the precipitate (likely barium or calcium) is to be separated by filtration, and the filtrate tested with sodium phosphate. This test is characteristic of magnesium.

ZINC.

Symbol, Zn. Atomic weight, 64.9.

Zinc has been indirectly known since the earliest days, as a substance that modified the appearance of copper, *brass* being the form produced through its agency. Basil Valentine first used the term (zinken). Zinc was once confounded with bismuth. The metal occurs in many localities in ores of various compositions, and is purified in large amounts and employed both in the arts and in medicine. Formerly calamine (an impure form of zinc ore, $ZnCO_3$) was the only ore worked, but at present many others are employed in this country and in Europe. Zinc is a blueish-white metal, melts at 433° C. and boils at 1040° C. The specific gravity is 6.9. It may be powdered at a temperature of about 205° C., but is malleable at between 100° and 150° C. (212° and 302° F.) The pure metal is scarcely affected by dilute sulphuric acid, but the addition of a few drops of solution of chloride of platinum causes immediate action, explained by the fact that the chlorine of the platinum salt unites with the zinc to form chloride of zinc, and the platinum is thus deposited upon the surface of the zinc, and induces a galvanic current. Commercial zinc is very impure. It contains in varying proportions lead, silver, cadmium, iron, arsenic, and even other metals, and carbon. It dissolves quickly in diluted sulphuric acid. Ordinary zinc is used in making brass, in galvanizing iron, in the form of sheets, and in many branches of the arts, and is a valuable metal. When it is melted and poured in a small stream into cold water it assumes a form known as *granulated zinc*, which is used to dissolve in acids.

Zinc Acetate. — Formula, $Zn2C_2H_3O_2$. Molecular weight, 182.62. Crystallized—$Zn2C_2H_3O_2.2H_2O$. Molecular weight, 218.54.

Acetate of zinc is made by the following—

Experiment—Place any given amount of acetic acid (officinal) in a capacious porcelain capsule, and add to it dry carbonate of zinc until effervescence ceases, and the carbonate is in slight excess. Warm the solution and filter it at once, then place the filtrate in a cool locality that crystals may form, which must be dried by exposure to the air upon porous paper. The mother liquor from the crystals may be evaporated and again crystallized. About two parts of carbonate of zinc will be required for five parts of acetic acid.

Acetate of zinc crystallizes in pearly white plates, very small and silky, from a concentrated solution, but they may be made of considerable size. The taste is bitterish and nauseating, they smell of acetic acid, slightly effloresce in dry air, and when heated fuse, blacken, burst into a flame, and leave a residue of oxide of zinc. They dissolve in half their weight of boiling water, in about three parts of cold, and in their weight of boiling alcohol, specific gravity 0.835.

Zinc Bromide (Bromide of Zinc), $ZnBr_2$.

This is produced when bromine vapor is passed over red-hot metallic zinc. It may be readily made by digesting hydrobromic acid with pure metallic zinc, filtering the product, and evaporating the filtrate to dryness, occasionally adding a little hydrobromic acid during the evaporation. By this experiment zinc unites with the bromine, and hydrogen gas escapes, thus:—

$$Zn+2HBr=ZnBr_2+H_2.$$

Bromide of zinc corresponds nearly in properties to the chloride. It has a sweetish, styptic taste, is white, deliquesces in the air, and sublimes on heating, condensing in white needles.

Zinc Carbonate.—Formula, $ZnCO_3$. Molecular weight, 124.75.

Synonym.—Normal zinc carbonate.

This salt in a hydrated form is thrown down when a solution of bicarbonate of potassium is mixed with a solution of sulphate of zinc. It is not used in medicine. Calamine, or zinc spar is a native impure form of this compound. If instead of hydrogen potassium carbonate, normal carbonate of sodium be employed to precipitate the sulphate of zinc, basic zinc carbonates are produced. The composition of these varies with both the temperature of the solutions and the proportion of water present. From this reason the so-called carbonate of zinc of our pharmacopæia is a basic carbonate.

Zinci Carbonas Præcipitata, U. S. P.—Formula, $ZnCO_3,2Zn(OH)_2$. Molecular weight, 322.39. Or, $ZnCO_3(ZnO)_2 3H_2O$. Molecular weight, 340.35.

Synonyms.—Precipitated carbonate of zinc. Carbonate of zinc.

Preparation.—Dissolve twelve troy ounces each of sulphate of zinc and of carbonate of sodium, severally, in four pints of distilled water with the aid of heat; then mix the solutions, and having stirred the mixture well, set it aside that the precipitate may subside. Pour off the supernatant liquor, and having washed the precipitate with hot water until the washings are nearly tasteless, dry it with a gentle heat. As thus prepared the so-called carbonate of zinc is a tasteless white powder, insoluble in water, but soluble in dilute acids with evolution of carbon dioxide. When it is calcined, water and carbon dioxide escape and oxide of zinc remains.

Zinc Chloride.—Formula, $ZnCl_2$. Molecular weight, 135.64.

Synonyms.—Zinci chloridum. Chloride of Zinc.

This may be prepared by dissolving ordinary metallic zinc in hydrochloric acid, (specific gravity 1.130 if convenient), and after the reaction, adding to the solution a few pieces of metallic zinc and digesting this mixture at the

temperature of a water bath for 24 hours, and then filtering the solution through muslin. Freshly precipitated carbonate of zinc, (precipitated in the cold with sodium carbonate), is then to be mixed with the solution and chlorine gas passed through it until in excess. The solution is now to be again filtered, and evaporated over a sand bath with constant stirring until a dry salt remains. The object of digesting the solution with excess of zinc is to precipitate lead, copper, arsenic and tin. The treatment with carbonate of zinc and chlorine separates the iron. Chloride of zinc is white, odorless, and possesses a disagreeable, saline, metallic taste. It is very deliquescent, and dissolves in both water and alcohol. It is used as a caustic and is employed in solution as an injection for preserving anatomical subjects. In this latter instance an impure solution is made by simply dissolving zinc in hydrochloric acid and straining..

Zinc Cyanide. (Cyanide of Zinc).—Formula, $Zn(CN)_2$. Or, $Zn(Cy)_2$.

This is made by adding hydrocyanic acid to a solution of acetate of zinc, cyanide of zinc precipitating and acetic acid being liberated. The precipitated cyanide of zinc dissolves to an extent in acetic acid, therefore, it is best to occasionally neutralize the most of the free acid with a little ammonia water. Cyanide of zinc is a snow-white powder, insoluble in water and alcohol, but dissolves in solution of cyanide of ammonium or cyanide of potassium. An excess of ammonia must not be used to neutralize the liberated acetic acid during the previously named operation. It decomposes by age.

Zinc Ferrocyanide.—Formula, Zn_2FeCy_6.

This is prepared by adding a solution of ferrocyanide of potassium to a solution of sulphate of zinc, and washing the precipitate well with water, and then drying it. Ferrocyanide of zinc is a white powder, and evolves hydrocyanic acid when added to dilute sulphuric acid. It is not soluble in water or alcohol.

Zinc Hydroxide. (Hydrate of Zinc). — Formula, $Zn(OH)_2$.

This is produced when caustic potash or caustic soda is added to the solution of a zinc salt. It is very soluble in an excess of the alkali, therefore, caution must be observed, that the proportions are nearly exact. According to Hodeker, the solution of this compound in caustic potash solution deposits crystals of the composition $Zn(OH)_2.H_2O$.

Zinc Iodide. (Iodide of Zinc).—Formula, ZnI_2.

Mix two troy ounces of iodine with about one troy ounce of granulated zinc, and eight fluid ounces of water, and digest together until the color disappears, towards the last, with the aid of a gentle heat. When the solution becomes transparent, filter it, then evaporate it carefully to dryness. Iodide of zinc is colorless, but upon exposure it deliquesces, then decomposes, and turns red, iodine being liberated. It dissolves freely in water, and a syrup of iodide of zinc is occasionally used in medicine. It may be made as follows:—

Syrup of Iodide of Zinc.—Filter the solution of iodide of zinc produced by the above formula, and add the filtrate to sixteen troy ounces of white sugar, and add to the mixture sufficient distilled water to bring to the measure of twenty fluid ounces. Dissolve the sugar by stirring.

Zinc Oxide. Formula, ZnO. Molecular wt. 80.86.

Synonyms. Flowers of zinc. Oxide of zinc.

Oxide of zinc is found native, [illegible] zinc ore [illegible] to [illegible] oxide [illegible] zinc [illegible] ceases [illegible] passing [illegible] oxide [illegible] Or by [illegible] oxide of [illegible] wish [illegible] assumes [illegible] cool-

ing. It dissolves in most dilute acids, and the solution should not strike a red color with sulphocyanide of potassium (iron), and should remain transparent upon treatment with sulphide of hydrogen (absence of lead, copper, etc.) Commercial oxide of zinc is usually a heavy, almost pure white powder, and is produced directly from the metal according to processes we have given. It is not desirable as a medicine.

Zinc Lactate. (Lactate of Zinc).—Formula, $ZnC_6H_{10}O_6.3H_2O$.

This salt is easily prepared by dissolving carbonate or pure oxide of zinc in dilute lactic acid, and evaporating the solution to crystallization. The crystals are acid in reaction, impart a metallic taste, and are quite soluble in hot water, but much less soluble in cold water.

Zinc Phosphide. (Phosphide of zinc), Formula—Zn_3P_2.

When finely divided metallic zinc is heated in the vapor of phosphorus, a gray mass results, having the above composition. It may be made of a bright metallic appearance by heating zinc to the melting point in an atmosphere of hydrogen, and then passing over it a mixture of hydrogen gas and phosphorus vapor until the phosphorus ceases to be absorbed. This compound can not be produced with safety in experimental amounts.

Zinc Salicylate. (Salicylate of zinc). Crystallized—$Zn2C_7H_5O_3.3H_2O$.

This salt is made by mixing salicylic acid with distilled water, heating the solution and adding an excess of oxide of zinc and digesting the mixture until the acid is saturated; then filtering the solution and crystallizing. The crystals are in the form of long needles, of a sweetish metallic taste, soluble in both water and alcohol.

Zinc Sulphate.—Formula, $ZnSO_4$. Molecular weight, 160 72. Crystallized—$ZnSO_47H_2O$. Molecular weight, 286.44.

Synonyms.—Zinci sulphas. Sulphate of zinc. White vitriol.

This salt, under the name of *white vitriol*, in contradistinction to *green* vitriol (sulphate of iron), and *blue* vitriol (sulphate of copper), has been known since the day of Basil Valentine. It was not, however, until 1735, demonstrated to be a compound of zinc. It is made on a large scale by roasting the native sulphide of zinc ores, and exhausting the calcined mass with water, and evaporating to crystallization. In commerce it is also made by dissolving metallic scrap zinc in dilute sulphuric acid, digesting for several hours at 180° F. the liquid in contact with excess of zinc, then filtering it and crystallizing. Sulphate of zinc usually contains traces of iron as made by this process, but the arsenic mostly escapes with the generated hydrogen, and what remains in solution precipitates upon subsequently digesting the solution with zinc, as is also the case with lead, cadmium, copper and tin. Iron may be separated by the process employed in making chloride of zinc. Sulphate of zinc may also be made by dissolving officinal carbonate or oxide of zinc in dilute sulphuric acid, and evaporating the solution to crystallization.

Sulphate of zinc is the most common salt of zinc, and is so cheap as to forbid its preparation by pharmacists excepting, simply, by way of an experiment. The crystals are isomorphous with Epsom salt (sulphate of magnesium), and in general appearance can not be distinguished from that article. Sulphate of zinc, however, will give a precipitate with ferrocyanide of potassium, and also with sulphide of ammonium, which is not the case with Epsom salt. Heated, it melts, and at 100° C. (212° F.) gives off six molecules of water, the remaining molecule being retained until a red heat is reached; at a white heat decomposition results and oxide of zinc remains. Sulphate of zinc is insoluble in alcohol, but it dissolves in two and one-third parts of cold water and in less than its own weight of boiling water. It has a disagreeable styptic taste, and in over doses acts as an emetic.

Zinc Sulphocarbolate. (Sulphocarbolate of zinc. Sulphophenate of zinc).—Formula, $Zn(C_6H_5SO_4)_2.8H_2O$.

This salt is made by mixing together equal weights of carbolic and sulphuric acids, and heating the mixture at a temperature of 180° F. for from six to ten hours, then diluting the mixture with eight parts of water and saturating the mixture with officinal carbonate of zinc, filtering the solution and evaporating it to crystallization. The crystals should be purified from a little adhering sulphate of zinc by recrystallization. This will form the ortho-salt. Para-sulphocarbolate of zinc is made in the same manner, excepting that the mixed sulphuric and carbolic acids are heated to 300° F. This latter salt crystallizes much nicer than the ortho-salt and is the modification found in market. Both forms have the same composition. Sulpho-carbolate of zinc should have a white color. It is inodorous, and dissolves readily in water.

Zinc Sulphide. (Sulphide of zinc).—Formula, ZnS.

This compound is found native, and the ore is known under the name, *zinc blend.* It usually contains iron and other impurities. It may be made by adding sulphide of ammonium to a solution of sulphate or chloride of zinc, washing the white precipitate with water and then drying it. Thus we obtain sulphide of zinc as an amorphous white powder, odorless and tasteless. When oxide of zinc is ignited in sulphide of hydrogen, crystals of sulphide of zinc form.

Zinc Valerianate.—Formula, $Zn2C_5H_9O_2.H_2O$. Molecular weight, 284.4.

Synonyms.—Zinci valerianas. Valerianate of zinc.

This is made by the officinal process by mixing a concentrated solution of valerianate of sodium with a strong solution of sulphate of zinc. Double decomposition follows, sulphate of sodium and valerianate of zinc being formed. The sulphate of sodium remains in solution while the valerianate of zinc crystallizes and may be washed with a little water and then dried. The reaction is expressed as follows:—

$$ZnSO_4 + 2NaC_5H_9O_2 = Na_2SO_4 + Zn2C_5H_9O_2.$$
Zinc sulphate. Sodium valerianate. Sodium sulphate. Zinc valerianate.

Valerianate of zinc may be made by dissolving pure oxide (or better, carbonate freshly precipitated) of zinc in dilute valerianic acid. When the acid is saturated, filter and evaporate the solution to crystallization, and then dry the crystals.. This salt is never made on a small scale, as the expense of valerianic acid, and loss of zinc valerianate forbids economy.

Valerianate of zinc crystallizes in white, light, mother-of-pearl or satin-like scaly crystals. The taste is bitterish and disagreeable; the odor is that of valerianic acid. When heated it decomposes, turns black, then bursts into flame and burns, leaving a residue of oxide of zinc. The crystals dissolve in cold water, in alcohol, and in ammonia water.

Tests for Zinc.—Ignite a compound of zinc on charcoal in the blow-pipe flame, and moisten the mass while hot with a solution of cobalt nitrate. Ignite again, a green mass (Rinmann's green) results.

2. Ignite on charcoal, a zinc compound with sodium carbonate, in the inner flame of a blow-pipe. Zinc oxide results, which is yellow when heated and white when cold. Moisten this with solution of cobalt nitrate, heat again, a green mass results.

3. Add ammonium sulphide to the neutral solution of a zinc salt; a white precipitate results. Add now solution of caustic potash until it is in excess; the precipitate does not dissolve. This reaction distinguishes zinc from all other metals.

LEAD GROUP OF METALS.

LEAD, THALLIUM.

These elements are very heavy, and comparatively soft. Lead in many respects resembles the metals of the alkaline earths, while thallium approaches the alkali metals. The basic oxides of each are alkaline. Their sulphides are black and insoluble.

LEAD (*Plumbum*).

Symbol, Pb. Atomic weight, 206.4.

Lead was mentioned in the Old Testament. It is one of the most abundant of metals, and has been found in a free state, but this is not common. The common ore of lead is *galena*, or lead sulphide, PbS, a mineral found in all parts of the world, and from this is produced the lead supply of the world. Lead is easily reduced from this ore. Roasting of galena in a reverberatory furnace at first changes a part of the sulphide of lead into sulphate of lead, by absorption of oxygen. An increase of heat now produces metallic lead and sulphur dioxide, thus:—

$$\underset{\text{Lead sulphate.}}{PbSO_4} + \underset{\text{Lead sulphide.}}{PbS} = \underset{\text{Lead.}}{2Pb.} + \underset{\text{Sulphur dioxide.}}{2SO_2}.$$

Commercial lead usually contains silver, antimony, copper and iron, and may contain nickel, zinc and bismuth. These impurities are of little consequence, however, as they are seldom in large proportions. Antimony renders it hard.

Lead melts at 334° C., is a soft bluish metal, not elastic, and is volatile at a white heat. It may be rolled into sheets, and becomes covered with a film of oxide upon exposure to moist atmosphere, and rapidly oxidizes when heated in a current of air. Dilute nitric acid dissolves it, but other dilute acids, as a rule, scarcely affect it. Metallic lead is extensively employed in the arts, and on account of its softness, low melting point, and capacity to resist the action of ordinary water and most dilute acids, is a valuable metal. The salts of lead are poisonous, and workmen engaged in white lead works, painters, etc., are subject to its effects.

Lead Acetate.—Formula, $Pb2C_2H_3O_2.3H_2O$. Molecular weight, 378 0.

Synonyms.—Plumbi acetas. Acetate of lead. Sugar of lead.

Experiment.—Mix two parts of officinal acetic acid with one part distilled water in an evaporating dish, and add to the mixture one and one-half parts of oxide of lead

(litharge). Heat the mixture on a water bath, with occasional stirring, and when the acid is neutralized filter it. Evaporate the filtrate until a pellicle forms, add a little acetic acid, place the solution in a cool situation to crystallize. Drain the crystals and dry them on paper in a cool situation.

As thus prepared, sugar of lead is obtained in the form of transparent needle-like crystals, but when made on a large scale it crystallizes in masses. Upon exposure it effloresces, and ultimately becomes, in part, converted into carbonate of lead, from absorption of carbon-dioxide from the air; therefore, it is often necessary to add a little acetic acid in dissolving acetate of lead. It is soluble in about two parts of cold water, in half its weight of boiling water, slightly soluble in ordinary alcohol, but is insoluble in ether. If it dissolves completely in water slightly acidulated with acetic acid (absence of traces of sulphate and chloride of lead), and the solution is free from blue tinge (copper), and has not the odor of coal gas, it may be considered pure enough for ordinary purposes. Zinc may be present, and even calcium and barium, but they are not likely to occur in large amount. Sugar of lead is an active poison.

Treatment.—Give solution of sulphate of magnesium (Epsom salt), or sulphate of sodium (Glauber's salt), or alum, or sulphate of potassium, and if vomiting does not take place, give an emetic of sulphate of zinc, or use the stomach pump. Let the patient drink freely of milk, and unless Epsom salt has been administered, give castor oil after the stomach has been evacuated by vomiting. For treatment of chronic cases of lead poisoning see treatises on poisons.

Solution of Subacetate of Lead is officinal in our pharmacopœia under the name *liquor plumbi subacetatis*. This was originally made by Goulard, and the common name for the solution at present is *Goulard's Extract*. It forms the preparation known as *Goulard's Water* when it is mixed with eighty parts of distilled water. This solution is not a definite compound, but a mixture of basic acetates. According to the U. S. P. Goulard's Extract

is made as follows: "Take of acetate of lead sixteen troy ounces; oxide of lead, in fine powder, nine troy ounces and a half; boiling water a sufficient quantity. Put the acetate and oxide into four pints of boiling water, in a glass or porcelain vessel, and boil for half an hour, occasionally adding boiling water to preserve the measure; then filter through paper. Lastly, keep the liquid in a well-stopped bottle."

Goulard's Extract is transparent, sweetish, astringent, and alkaline to red litmus paper. It precipitates upon exposure to the air, or upon mixing with ordinary water (carbonate of lead), but mixes transparent with pure distilled water. All the lead may be precipitated with sulphuric acid. It is poisonous, and the treatment is that given for poisoning by acetate of lead.

Normal Lead Carbonate.—Formula, $PbCO_3$. Molecular weight, 266.25.

This is a mineral known as *cerusite*. It is found native in the form of transparent crystals. It may be made by passing carbon dioxide into a weak solution of sugar of lead, when it will precipitate. This is not used in medicine, and is not the ordinary compound known as *carbonate of lead*. Lead gives rise to several basic carbonates, and the officinal carbonate of lead (white lead) is a *basic* carbonate.

Lead Carbonate.—Formula, $2PbCO_3+Pb(OH)_2$. Molecular weight, 772.82.

Synonyms.—Plumbi carbonas. White lead.

This is a very old compound, and strange as it may seem, the original "Dutch process," for its manufacture is preferred at present. This consists in placing rows of earthen pots under sheds, pouring into each pot a little acetic acid, placing above this, inside of each pot, a roll of sheet lead, and covering the entire arrangement with tan-bark. The heat of the decomposing bark evaporates the vinegar, this unites with the lead to form basic acetate

of lead, and the carbon dioxide supplied from the decomposing organic matter of the bark converts this into *white lead.* Several improvements have been suggested over this crude method, and in Europe some patents have even been taken, but it is doubtful if any process gives a white lead with the covering (body) power of the original. In this city we know preference is given to the original article.

Common carbonate of lead is white, heavy, and amorphous, even under the microscope showing no crystalline structure. It should dissolve in dilute nitric and acetic acids without residue. It is turned black by sulphide of hydrogen, and from this reason is not admired as a paint for laboratories or the buildings near sulphur springs. It is adulterated by "baryta" (sulphate of barium) to a very great extent, but this adulteration is admitted by such as purchase cheap brands of white lead. In this country the manufacturer's word is now considered a guarantee of the quality. Sulphate of barium will not dissolve in dilute nitric acid, therefore the adulteration is easily detected. White lead is poisonous, and for the treatment we refer the reader to *acetate of lead*, page 254.

Lead Chloride.—Formula, $PbCl_2$.

Synonym.—Chloride of lead.

Lead will slowly combine with chlorine at ordinary temperatures. Boiling hydrochloric acid with lead liberates hydrogen and forms chloride of lead. Solution of alkaline chlorides form precipitates of lead chloride with solutions of lead salts. Hydrochloric acid acts on carbonate or oxide of lead, chloride of lead resulting. There are several basic chlorides of lead (oxy-chlorides) some of which are native, as the minerals mallockite and mendipite. A form of oxy-chloride of lead, Pb(OH)Cl, is used as a substitute for white lead in some countries, and is made by a patented process. *Turner's yellow* is an oxychloride of lead. None of the compounds of chlorine with lead are used in medicine. The antidote for lead chloride is the same as that for acetate of lead.

Lead Iodide.—Formula, PbI_2. Molecular weight, 459.46.

Synonyms.—Plumbi iodidum. Iodide of lead.

This may be made by dissolving metallic lead in hot hydriodic acid, but the easiest process is that adopted by the pharmacopœia, essentially as follows :—

Experiment.—Dissolve one part of iodide of potassium in two parts of distilled water, and mix this with a solution of one part of nitrate of lead in six parts of distilled water. Filter the liquid, and wash the precipitate with a little cold distilled water, then dry it on bibulous paper by exposure to the atmosphere. Iodide of lead is a beautiful yellow compound, soluble in 1235 parts of cold water and 194 parts of boiling water (R. and S.), the solutions being colorless. The specific gravity is 6.1. It dissolves freely in hot solution of chloride of ammonium, and an interesting experiment is to dissolve it in this manner in a beaker glass, and observe the beautiful play of colors as the light is reflected from the glittering scale-like crystals during their formation. This is especially beautiful in sunlight.

Lead Nitrate.—Formula, $Pb(NO_3)_2$. Molecular weight, 330.18.

Synonyms.—Plumbi nitras. Nitrate of lead. [Lead saltpetre.] Normal lead nitrate.

Nitric acid, warm, acts upon lead, the result being nitrate of lead. This salt is manufactured upon a large scale by dissolving litharge in dilute nitric acid by the aid of heat, and evaporating the solution to crystallization. Nitrate of lead crystallizes in white porcelain-appearing crystals of an astringent, metallic taste. It dissolves freely in water, with decrease of temperature and deflagrates upon rubbing with sulphur. This salt is largely used in the arts, as in dyeing, and is used in medicine, and in the preparation of iodide of lead. It forms, when boiled with water and its weight of litharge, a compound, $Pb(NO_3)OH$, which is a basic nitrate of lead. The antidote for nitrate of lead is that suggested for acetate of lead.

Lead Oleate. — Formula, $Pb2C_{18}H_{33}O_2$. Molecular weight, 767.16.

Synonyms.—Oleate of lead. Lead plaster.

When litharge is boiled with olive oil, decomposition ensues and oleate of lead and glycerin results. This compound has long been familiar, under the name *lead plaster*, and in addition numbers of ointments and unofficinal plasters contain oleate of lead as a base, either litharge or red lead being boiled with various oils in their preparation. Among such we name Mayer's ointment, black salve, etc. Oleate of lead has been employed in medicine and is prepared from oleic acid as follows: Place one hundred parts of oleic acid in an evaporating vessel, and mix with it twenty parts of pure litharge. Heat the mixture to a temperature of 180° F., with stirring, maintaining the heat until the litharge is dissolved. Oleate of lead, as thus prepared, is a tenacious, sticky mass. In former years the formation of lead plaster was used as a source of glycerin.

Lead Oxide.—Formula, PbO. Molecular weight, 222.36.

Synonyms.—Plumbi oxidum. Oxide of lead. Lead monoxide. Litharge.

This compound exists naturally as the mineral *Noggerath.* It is formed when lead is simply heated in the air. and has been known since ancient times. Litharge is made by heating metallic lead in a draught of air, and scraping from its surface the layers of oxide as they form. This contains metallic lead, has a yellowish-gray color, and is known as massicot. It is now more strongly heated in the atmosphere whereby the lead is perfectly oxidized and the mass partly fuses. The color of litharge may incline from yellow to reddish. Litharge is often impure, containing iron, copper and carbonate of lead in small proportion. It is largely employed in the arts.

Red Lead, Pb_3O_4, is either made from litharge or carbonate of lead, by careful roasting at dull redness. There are other oxides of lead, but they are of no particular importance to physicians or pharmacists.

Lead Sulphate.—Formula, $PbSO_4$. Molecular weight, 302.22.

Synonyms.—Plumbi sulphas. Sulphate of lead.

Sulphate of lead is found native and known as *lead vitriol.* It is easily made by mixing the solution of an alkaline sulphate with a solution of a soluble lead salt, the sulphate of lead precipitating as a white powder. It has the specific gravity of about 6.2, is almost insoluble in water, but dissolves in solution of many salts, such as acetate of ammonium and acetate of potassium. It is of no particular importance.

Lead Sulphide.—Formula, PbS.

This compound occurs native as *galena,* and is the natural source of lead and its compounds, therefore it is very valuable, but is not used in the arts nor in medicine. Sulphide of lead is thrown down as a black precipitate when sulphide of hydrogen is passed through a solution of a salt of lead, and as we have said, is formed when sulphide of hydrogen comes in contact with white lead paint.

Tannate of Lead is made by pouring a solution of tannic acid into a solution of nitrate of lead. It precipitates nearly white, but when the precipitate is washed and dried it changes to a brown color.

Tests for Lead.—1. Pass sulphide of hydrogen through a solution of a lead salt, black sulphide of lead is produced. If some other metals such as mercury and copper are present, the reaction is identical, as the sulphides are also insoluble and black. To identify lead, mix the precipitate with warm nitric acid, filter, add to the filtrate a little sulphuric acid; a white precipitate will follow if lead was present in the original solution.

2. If hydrochloric acid be added to a solution of a lead salt (not too dilute) a white precipitate results. This dissolves in boiling water and separates when the solution cools, as needle-like crystals. If the solution of the lead salt is very dilute, the precipitate fails to form. In such cases perform the sulphide of hydrogen test.

THALLIUM.

Symbol, Tl. (T). Atomic weight, 203.6.

Thallium is one of the recent discoveries, having been identified by spectrum analysis in 1861, by Crookes. He supposed it to be of the sulphur group, and considered it a non-metallic element. In 1862, however, Mr. Lamy, independent of Mr. Crooke's discovery, prepared the element in metallic form.

Thallium exists in many varieties of iron and copper pyrites and in a few other minerals, and it is found in some natural waters. The metal has a bluish tint, is very soft, malleable, has the specific gravity 11.8, and melts at the temperature of 290° C. It dissolves in dilute acids, forms two oxides (monoxide Tl_2O, and trioxide Tl_2O_3) and a hydroxide, TlOH. It forms salts, and is peculiar in this respect, that it resembles the alkali metals in some particulars, as in its alkaline hydroxide. On the other hand, it resembles lead as an element, and its sulphide is black. It is of no use in the arts, and none of its salts are employed in medicine.

THE COPPER GROUP OF METALS.

COPPER, MERCURY, SILVER.

COPPER. (*Cuprum*).

Symbol, Cu. Atomic weight, 63.0

This metal is found native, and has been used by man from pre-historic times. It occurs in many localities, sometimes, as near lake Superior, in enormous masses. It usually contains small amounts of silver, bismuth, lead, iron, etc. It is unnecessary for us to go into the history of copper smelting, and other processes for extraction of the metal. Copper is sometimes found native in crystalline form, as octohedrons. Copper has a reddish color, known as "copper-red," and when polished acquires a bright metallic luster. It is very tough and malleable,

but the admixture of certain other metals, and of sulphur, renders it brittle. Excepting silver, it is the best known conductor of electricity. It melts at a temperature less than the melting point of gold and greater than that of silver. When molten it has a greenish-blue color and possesses the power of absorbing many gases.

Copper forms two series of salts, known as *cuprous* and *cupric* salts. These correspond to the oxides, thus:—

{ Cuprous oxide.....Cu_2O.
{ Cuprous bromide.....Cu_2Br_2.
{ Cupric oxide.....CuO.
{ Cupric bromide.....$CuBr_2$.

Of these two classes of salts, only the *cupric* salts are of use in medicine, therefore we shall not consume space by a study of the *cuprous* salts.

Alloys of Copper.—Brass is an alloy of copper and zinc. Various qualities of brass are made, containing different proportions of the elements, and as the proportion of copper to zinc increases, the color of the alloy reddens. Copper-zinc alloys are harder than copper. Occasionally small amounts of other metals are added.

Antidotes.—The salts of copper are poisonous. Usually an overdose excites violent vomiting, and if this is not the case, either an emetic or the stomach pump must be employed. Mucilaginous drinks, milk and raw eggs may be given with advantage, and fine iron filings, (see sulphate of copper). Copper poisoning, it is said, may result from the copper in pickles, and even preserved fruits, and in such cases the copper is derived from the vessel in which they were prepared.

Copper Acetate.—Formula, $Cu2C_2H_3O_2$. Molecular weight, 180.72.

Synonyms.—Acetate of copper. Cupri acetas. Cupric acetate.

This is made by dissolving verdigris (sub-acetate of copper) in acetic acid in excess, adding a little water, filtering and evaporating the filtrate to crystallization. Crystals of acetate of copper are small, dark green, or inclined to bluish, of a metallic taste and soluble in water.

This salt is not used in medicine, but the solution is known in commerce as—

Tinctura Cupri Acctici Rademacheri, or Rademacher's tincture of acetate of copper. It is made by dissolving one part of acetate of copper in crystals in ten parts of warm water, and adding to the solution eight parts of alcohol. The original process of Rademacher was by the double decomposition of crystallized sulphate of copper and crystallized acetate of lead, and we prefer it at the present day. It is as follows:—

Take of crystallized sulphate of copper, and of crystallized acetate of lead, each three (3) troy ounces; water, seventeen (17) fluid ounces; alcohol, thirteen (13) fluid ounces. Powder the the salts separately and then rub them together in a mortar until they assume a pasty condition.† Then gradually add the water, and afterward the alcohol. Place the mixture in a bottle, and after fourteen days, carefully decant the clear dark green or blue solution from the precipitate of lead sulphate, and filter it through paper.

Copper Sub-acetate.—Formula, $(CuO)2C_2H_3O_2$. Molecular weight, 196.68.

Synonyms.—Cupri subacetas. Sub acetate of copper. Oxy-acetate of copper. Cupric sub-acetate. Verdigris.

This compound is an article of commerce in a more or less impure state. It is usually employed as a constituent of horse liniments, and for this reason the article is seldom required pure. Verdigris is made in countries where wine is produced, the manner of preparation being as follows: Grape husks are permitted to ferment, and during the action placed in layers in contact with partly corroded

† This, and the process for making tincture of acetate of iron, illustrate in a beautiful manner the liberation of water of crystallization. In this instance the reaction is explained as follows:—

$$CuSO_4.5H_2O+Pb2C_2H_3O_2.3H_2O=$$
$$PbSO_4+Cu2C_2H_3O_2+8H_2O.$$

It will be observed that the combined water of the two original salts is equal to eight molecules, all of which are liberated, for the resultant salts are anhydrous.

copper plates. This forms acetate of copper, which is changed into sub-acetate by subsequent wetting and exposure to the atmosphere. Verdigris is an impure sub-acetate of copper, containing oxide of copper. It should mostly dissolve in ammonia, and either diluted sulphuric, hydrochloric, or acetic acid. Effervescence should not take place while making the solutions in the acids.

Copper Arsenite, $CuHAsO_3$, or *Scheele's green*, is made by precipitating a solution of arsenite of potassium with a solution of sulphate of copper. This was originally prepared by *Scheele* and from him it derived the well-known and common name. It is of a beautiful green color and is extensively employed as a pigment both in painting and in paper coloring. It is very poisonous, although it is insoluble in water, for the acids of the stomach render it soluble. The antidote is calcined magnesia, freshly precipitated ferric oxide, and albumen. Use emetics and the stomach pump (see antidotes for arsenious acid). Scheele's green is not used in medicine. There are several native basic *arseniates* of copper, and they can also be made artificially, but are of no interest as connected with our subjects.

Copper Carbonate. (Carbonate of copper.)

This compound, known also as sub-carbonate of copper and basic carbonate of copper, is the compound formed when sulphate of copper solution is precipitated by solution of sodium carbonate. It is of a greenish color, and when freshly precipitated has the composition $CuCO_3 + Cu(OH)_2 + H_2O$. Upon drying the precipitate the water escapes and a salt of the composition $CuCO_3.Cu(OH)_2$ remains. The so-called copper rust which forms on metallic copper by exposure to the air (sometimes called verdigris) is this compound. Carbonate of copper prepared as above, dissolves perfectly with a green color in dilute acids and with a blue color in ammonia water. The most likely impurity will be sulphate of sodium which can be detected by boiling the carbonate with water, filtering,

and adding solution of barium chloride to the filtrate; a white precipitate will show a salt of sulphuric acid. There are other basic carbonates of copper, but none of use in pharmacy or in medicine.

Copper Hydroxide, or *Cupric hydrate,* $Cu(OH)_2$, is made when a solution of sulphate of copper is precipitated with the solution of an alkali. It has a blue color, and dissolves in ammonia to form a deep blue liquid, and to this fact we are indebted for one of the tests for copper. This solution will also dissolve cotton, linen, and other modifications of cellulose.

Copper Oxide.

There are four oxides of copper, one only requiring notice from us.

Copper Monoxide, Cupric oxide, Oxide of copper, (CuO), or black oxide of copper. This is found native as the mineral *melaconite,* but is usually made by calcining the carbonate of copper. For this purpose the well washed and dried carbonate of copper is pressed into a Hessian crucible, and with a gradually increasing heat raised to redness, and maintained at that temperature half an hour. When nearly cool the black powder is placed in glass stoppered vials.

Black oxide of copper is tasteless and odorless. It dissolves in dilute acids, and the addition of ammonia to such solutions produces deep blue liquids, when the ammonia is in excess, and greenish precipitates when the ammonia is in small amount.

Copper Sulphate.—Formula. $CuSO_4$. Molecular weight, 158 82. Crystallized—$CuSO_4 5H_2O$. Molecular weight, 248.62.

Synonyms.—Cupri sulphas. Sulphate of copper. Cupric sulphate. Blue vitriol.

This salt may be made by mixing water and sulphuric and nitric acids, then boiling the mixture in contact with the copper until all action ceases, and then crystallizing the salt. Or it may be made by dissolving in the same manner copper scales obtained from the copper-

smith. The sulphate of copper of commerce is obtained as an incidental product in certain branches of the arts, such as silver refining, or is produced from copper ores or as a waste from copper mines, and as a consequence it can not be made directly with a pecuniary object in a small way. Sulphate of copper forms transparent, blue crystals, of the specific gravity 2.28. It has no odor and the taste is metallic and astringent The crystals effloresce upon the surface in dry air, and turn white. At a temperature of 240° C. it parts with the water of crystallization, and forms a white anhydrid, which absorbs water from the air and becomes blue. At a red heat decomposition ensues and black oxide of copper remains. Sulphate of copper dissolves in about its weight of boiling water, and in about three and one-half parts of cold water. It is an emetic. The solution of sulphate or other cupric salt deposits metallic copper upon a bright blade of iron or steel, and this fact has suggested the use of iron filings as an antidote for poisoning by copper salts. Sulphate of copper forms a compound with ammonia known as—

Ammoniated Copper. (Ammoniated sulphate of copper.)—Formula. $CuSO_4.4NH_3.H_2O$.

This may be made by rubbing powdered sulphate of copper with ammonia water until the sulphate dissolves, and then mixing the solution with alcohol, and collecting the precipitate which forms, and drying it quickly. This salt decomposes upon long exposure to the air, the result being a basic sulphate of copper and sulphate of ammonium. Ammoniated copper dissolves in one and one-half parts of cold water, but decomposes in a large amount of water (Wittstein). This salt is occasionally used in medicine, and is made by rubbing together one part of sulphate of copper and one and one-half parts of officinal carbonate of ammonium, and drying the pasty mass quickly. As thus obtained it is a variable compound.

Tests for Copper.—Sulphide of hydrogen throws down black sulphide of copper from solutions of the copper salts. This precipitate dissolves in warm nitric acid, the solution having a green color. The addition of ammonia

in excess changes the color to dark blue. When ammonia water is slowly added to the solution of a copper salt a green precipitate forms at first, but this redissolves when the ammonium reaches an excess, a dark blue liquid resulting. When solution of potassium ferrocyanide is added to the solution of a cupric salt, a red brown precipitate results. This precipitate is insoluble in dilute acids, but dissolves in ammonia water. If a strip of bright iron be introduced into the solution of a copper salt the metal will be deposited upon its surface.

MERCURY. (*Hydrargyrum*).

Symbol, Hg. Atomic weight, 199.8.

Mercury is also known as *quicksilver*. It has been known since the day of the alchemists, and at one period they believed it to be the fundamental principle of all metals. Mercury is found in many parts of the world as an oxide, *cinnabar*. Most extensive mines are in Spain and California.

Mercury is a very interesting substance, standing isolated among metals, for it is fluid at ordinary temperatures. It resembles silver in appearance, and is so heavy that iron and stone float upon its surface, as cork upon water. Commercial mercury is usually impure, and when a globule of it is permitted to flow over a clean surface, a "tail" remains, caused by the presence of dissolved metals. Mercury may be purified by distillation, the surface of the metal, during the operation, being covered with iron filings. Pure mercury is fluid at ordinary temperatures. It is silver-white, and is said to be transparent when in thin films, transmitting a blue light. Its specific gravity at 0° C. (32°F.)is 13.595. Mercury freezes at —39.4° C. and boils at 357.25° C.

Compounds or mixtures of mercury with other metals are called *amalgams*. When the proportion of mercury is large, the amalgams are liquid. When the other metal is in large proportion the amalgam is pasty or granular. In combining with metals, heat may be either liberated, as with sodium, or absorbed, as with tin. Amalgams

are very unstable, if they really are true chemical compounds. It is customary to speak of *dissolving* metals in mercury. None of the amalgams are of interest in a pharmaceutical sense, but the principle of making them may be illustrated by rubbing a clean copper coin with a little metallic mercury, when a silver color will be produced. Metallic mercury is used extensively in the refining of silver and gold, in thermometers and like instruments, in silvering mirrors (amalgamated with tin), and in many other branches of the arts. Rubbed with confection of roses it forms—

Pilulæ Hydrargyri or *Pills of Mercury,* commonly known as *blue mass.* This is made as follows, according to the U. S. P.

Take of mercury, three hundred and eighty-five grains; confection of rose, five hundred and seventy-six grains; liquorice root, in fine powder, one hundred and ninety-two grains. Rub the mercury with the confection until the globules cease to be visible; then add the liquorice root, beat the whole into a pilular mass, and divide into three hundred and eighty-four pills. It is customary to prepare and sell this article to druggists in mass.

Unguentum Hydrargyri, or *Mercurial Ointment,* is made as follows:—

Take of mercury, twenty-four troy ounces; of lard and suet, each, twelve troy ounces. Rub the mercury with a troy ounce of the suet, and a small portion of the lard, until the globules cease to be visible; then add the remainder of the lard and of the suet, soften with a gentle heat, and mix them.

It will appear as though the properties of the above named preparations should resemble those of metallic mercury, but it has been demonstrated that metallic mercury has no action upon the animal economy when swallowed, even in large amounts. Some have argued that "extinguished" mercury (rubbed with other substances until globules are not visible to the eye) is in reality, partly an oxide, and thus account for its action. It seems to the writer, however, that the fine state of comminution, and

extent of surface produced, is sufficient to give the juices of the stomach and intestines an opportunity to act upon the metal which is impossible with a few large globules. Mercury in the state of vapor is poisonous, and many of its salts are very poisonous. It forms two oxides, Hg_2O, *mercurous* oxide; and HgO, *mercuric* oxide. It also forms two classes of salts, which correspond to the two oxides and are known as *mercurous* salts, and *mercuric* salts, thus:—Hg_2I_2 is the lowest iodide, while HgI_2 is the highest iodide. Both mercurous and mercuric salts are used in medicine.

Mercurous Chloride.—Formula, Hg_2Cl_2. Molecular weight, 470.34.

Synonyms.—Hydrargyri chloridum mite. Mild chloride of mercury. Calomel. Subchloride of mercury.

This compound may be made by direct combination of its elements. By acting upon a solution of mercurous nitrate with dilute hydrochloric acid, or with a solution of common salt (chloride of sodium). A very old process is to distill a mixture of mercury, mercuric sulphate and common salt. This last is the present officinal process. It is unnecessary for us to reproduce it in detail, as calomel is *never* made in a small way, and few desire to prepare it even for the sake of experiment.

Calomel is a white powder; when made by sublimation it is in larger particles than when made by precipitation, and many physicians claim that the precipitated article is preferable. It is tasteless, odorless, and evaporates at a red heat, but does not fuse. It is insoluble in water, alcohol, and ether; and distilled water boiled with it after filtration yields no precipitate with ammonia or nitrate of silver, (absence of soluble salts of mercury). With solution of caustic potash it turns nearly black, suboxide of mercury being formed. A preparation known as "black wash," is made by mixing calomel with lime water.

Mercurous Iodide.—Formula, Hg_2I_2. Molecular weight, 652.66.

Synonyms.—Hydrargyri iodidum viride. Green iodide of mercury. Protiodide of mercury.

This salt is made by rubbing together mercury and iodine, or by triturating mercuric iodide (red iodide of mercury) with metallic mercury, or by the action of solution of mercurous nitrate upon solution of iodide of potassium.

According to the U. S. P., one troy ounce of mercury is to be rubbed with three hundred grains of iodine, half an ounce of alcohol being added to the mixture. I prefer to place the *alcohol* in the mortar first. The result is to be washed with stronger alcohol until the washings cease to produce a cloudiness when dropped into water. The elements combine directly to form mercurous iodide, yet at the same time an uncertain amount of *mercuric* iodide is always formed, and this is separated by the subsequent treatment with alcohol, in which it is soluble. It is doubtful, however, if commercial mercurous iodide is ever free from mercuric iodide.

Green iodide of mercury is a greenish-yellow powder, insoluble in water and alcohol. When treated with stronger alcohol, the filtrate should scarcely cloud distilled water when dropped into it, and should leave but a faint red stain upon porcelain after evaporation. It sublimes as yellow crystals (Yvon) when heated to 250° C. Mercurous iodide becomes of a red color when heated.

Mercurous Nitrate.—Formula, $Hg_2(NO_3)_2$. Molecular weight, 523.38.

Synonym.—Nitrate of mercury.

When mercury is dissolved in cold nitric acid, especially if dilute, this compound is produced according to the following equation:—

$$\underset{\text{Mercury}}{2Hg_3} + \underset{\text{Nitric acid}}{8HNO_3} = \underset{\text{Mercurous nitrate.}}{3Hg_2(NO_3)_2} + \underset{\text{Water.}}{4H_2O} + \underset{\text{Nitric oxide.}}{2NO}$$

The chemistry of this salt is of little interest, in a pharmaceutical sense, excepting that its solution is produced as a preparatory step in making citrine ointment. However, in this case, the after action of the hot grease leaves it problematical as to the composition of the salt which enters into the constitution of the ointment. Often also, the action of the nitric acid upon the mercury, produces

heat sufficient to form a portion of *mercuric* nitrate, and perhaps where strong nitric acid is used, the *mercuric* nitrate is in largest proportion.

Mercurous nitrate is soluble in dilute nitric acid and in a small proportion of water, but an excess of water produces a basic salt. It crystallizes with two molecules of water of the monoclinic system of crystallization.

Mercurous Oxide.—Formula, Hg_2O. Molecular weight, 415 56.

Synonyms.—Sub-oxide of mercury. Black oxide of mercury.

This is obtained by acting upon mercurous salts with the caustic alkalies. It is seldom used in medicine, if we may except the so-called black-wash prescribed by physicians, and which is made by authority of the British pharmacopœia by mixing thirty grains of calomel with ten fluid ounces of lime-water. Mercurous oxide is of a blackish-brown color, and is easily decomposed by certain organic substances, into metallic mercury and mercuric oxide.

Mercuric Ammonium Chloride.—Formula, NH_2HgCl. Molecular weight, 251.18.

Synonyms.—Mercuric-ammonium chloride. Hydrargyrum ammoniatum. Ammoniated mercury. Infusible white precipitate. White precipitate.

This salt has the composition of ammonium chloride, NH_4Cl, excepting that two atoms of hydrogen are displaced by one atom of mercury. It is made by pouring a solution of corrosive sublimate into excess of ammonia water, the reaction being as follows:—

$$\underset{\text{Mercuric chloride.}}{HgCl_2} + \underset{\text{Ammonia}}{2NH_4OH} = \underset{\text{Ammoniated mercury.}}{NH_2HgCl} + \underset{\text{Ammonium chloride.}}{NH_4Cl} + \underset{\text{Water.}}{2H_2O}.$$

Experiment.—(U. S. P. proportions.) Dissolve six parts of corrosive sublimate in ninety-six parts of distilled water by aid of heat, cool the solution, and stir into it eight parts of ammonia water. Wash the precipitate by decantation with distilled water, then dry the precipitate in a cool dark place.

Ammoniated mercury is a white powder, insoluble in water and alcohol, but soluble in muriatic acid. Lime water will not blacken it. When heated it decomposes at less than redness *without fusing*, and is entirely dissipated, the products being calomel, nitrogen, and ammonia gas, thus:—

$$6NH_2HgCl = 3Hg_2Cl_2 + N_2 + 4NH_3.$$

Ammoniated mercury should not be mixed with iodine and alcohol, as this will produce a dangerous explosion. There are many other mercuric-ammonium compounds, none of interest from a pharmaceutical sense, however, excepting—

Mercuric-diammonium Chloride. (Fusible white precipitate.)—Formula, $(NH_3)_2HgCl_2$.

This may be made by boiling ammoniated mercury with solution of chloride of ammonium, or by precipitating solution of mercuric nitrate with sal-ammoniac and carbonate of potassium. At one time this compound was confused with ammoniated mercury (NH_2HgCl). Fusible white precipitate melts at a temperature less than redness.

Mercuric Chloride. — Formula, $HgCl_2$. Molecular weight, 270.54.

Synonyms.—Hydrargyri chloridum corrosivum. Corrosive chloride of mercury. Perchloride of mercury. Bichloride of mercury. Corrosive sublimate. Mercurius sublimatus corrosivus.

Bichloride of mercury is produced when mercury is heated in chlorine gas, but in practice it is made by subliming mixtures of sulphate of mercury (mercuric sulphate) and chloride of sodium (common salt), mixed with a small amount of manganese dioxide. The reaction may be expressed as follows:—

$$\underset{\text{Mercuric sulphate}}{HgSO_4} + \underset{\text{Sodium chloride.}}{2NaCl} = \underset{\text{Mercuric chloride.}}{HgCl_2} + \underset{\text{Sodium sulphate.}}{Na_2SO_4}.$$

Sometimes mercuric sulphate is contaminated with mercurous sulphate. In that case *calomel* will be formed and sublime with the corrosive sublimate. The addition

of a little black oxide of manganese (manganese dioxide) to the mixture previous to heating it, liberates a small amount of chlorine gas during the reaction, and this converts the calomel (Hg_2Cl_2) into corrosive sublimate $HgCl_2$.

Corrosive sublimate occurs in commerce as white masses, semi-transparent, presenting a crystalline fracture. It dissolves in twice its weight of boiling water, but is much less soluble in cold water. It dissolves in three parts of alcohol, four parts of ether, and crystallizes from solution in four-sided prisms. It is very heavy, having a specific gravity of 5.403. Heat melts and then volatilizes it. With many chlorides corrosive sublimate forms double chlorides, which are as a rule more soluble in water than corrosive sublimate. This fact is taken advantage of for making "bed-bug" poison, equal weights of corrosive sublimate and sal-ammoniac being dissolved in water for this purpose. Mercuric chlorides form a number of basic chlorides, which may be obtained by boiling a solution of corrosive sublimate with mercuric oxide. Corrosive sublimate has a sharp metallic taste and is an active poison. The best antidote is *albumen*, with which it forms an insoluble compound, and for this reason white of eggs should be given at once. If eggs can not be obtained milk may be freely administered, or linseed and warm water. In all cases vomiting must be provoked as soon as possible, by the use of emetics, or the stomach emptied with a stomach pump. All of the soluble, and many of the insoluble salts of mercury are active poisons. The general treatment is that which we have mentioned for corrosive sublimate.

Mercuric Cyanide.—Formula, $Hg(CN)_2$. Molecular weight, 251.76.

Synonyms.—Cyanide of mercury. Hydrargyri cyanidum.

This is made according to the U. S. P. by distilling a mixture of ferrocyanide of iron, sulphuric acid and water, and neutrallizing the distillate with red oxide of mercury and evaporating the product in a dark place to crystallization. It may be made more simply by dissolving red

oxide of mercury (red precipitate) in dilute hydrocyanic acid, filtering, and evaporating the filtrate to crystallization, the reaction being as follows:—

$$\underset{\text{Mercuric oxide.}}{HgO} + \underset{\text{Hydrocyanic acid.}}{2HCN} = \underset{\text{Cyanide of mercury.}}{Hg(CN)_2} + \underset{\text{Water.}}{H_2O}.$$

Scheele discovered this compound, obtaining it by boiling Prussian blue, water and red oxide of mercury together. Cyanide of mercury is insoluble in pure alcohol, but dissolves in eight parts of water. It crystallizes in needles, and when heated, decomposes, giving off cyanogen. When hydrochloric acid is added to the solution of cyanide of mercury, chloride of silver is formed as a white precipitate, and hydrocyanic acid is set free, thus:—

$$2HCl + Hg(CN)_2 = HgCl_2 + 2HCN.$$

Cyanide of silver is sometimes employed for the purpose of making hydrocyanic acid.

Mercuric Iodide.—Formula, HgI_2. Molecular weight, 452.86.

Synonyms.—Hydrargyri iodidum rubrum. Red iodide of mercury. Biniodide of mercury.

Experiment.—Dissolve one part of corrosive sublimate in thirty parts of water, and pour into this solution a solution of one-fourth part of iodide of potassium in eight parts of water. Wash the precipitate with distilled water, and dry it by exposure to the air on bibulous paper. When solution of iodide of potassium is slowly poured into solution of corrosive sublimate, mercuric iodide is formed, and chloride of potassium, thus:—

$$\underset{\text{Mercuric chloride.}}{HgCl_2} + \underset{\text{Potassium iodide.}}{2KI} = \underset{\text{Mercuric iodide.}}{HgI_2} + \underset{\text{Potassium chloride.}}{2KCl}.$$

It will be noticed that as the first portions of the solutions come together, a yellowish cloud forms, which immediately dissolves; after a little, it ceases to dissolve, turns red, and precipitates. This precipitate will dissolve in an excess of either solution, therefore, it is necessary to have the components in proper proportion. Red iodide

of mercury is a brilliant, deep-red or scarlet, crystalline powder, it turns yellow when heated to 150° C. (302° F.) and regains the red color upon cooling. If red-iodide of mercury be melted and then cooled, it remains yellow, but a blow from a hard body at once turns the mass to red. It dissolves in solution of chloride of sodium, in hot alcohol, in solution of corrosive sublimate or in solution of iodide of potassium; but is insoluble in water. It also dissolves in solutions of many ammonium salts and soluble iodides, forming double compounds. If red iodide of mercury be adulterated, the fact can be discovered by subliming it, or dissolving it in hot alcohol, when the impurities will remain. A specimen that once came under our notice nicely crystallized, contained a considerable amount of scaled glass. Red iodide of mercury is very poisonous, and great care must be taken that it is not mistaken for the green iodide, and substituted where that salt is intended. We can not name a reliable chemical antidote; the main dependence must rest with mucilaginous and albuminous drinks, emetics, and the stomach pump.

Mercuric Oxide.—Formula, HgO. Molecular weight, 215.76.

Synonyms.—Hydrargyri oxidum rubrum. Red oxide of mercury. Red precipitate. Mercuric monoxide.

This was originally obtained by heating mercury some length of time in the air. It is made on a large scale by heating a mixture of nitrate of mercury and metallic mercury until red fumes cease to be evolved. The U. S. P. directs that nitrate of mercury alone, be heated "until red vapors cease to be evolved." Few will care to make this article for the sake of experiment, and it can not be made on a small scale to any pecuniary advantage. Red oxide of mercury is a bright red crystalline powder, which in a strong heat turns black, then decomposes, evolving oxygen; metallic mercury remaining.

When solution of corrosive sublimate is poured into solution of caustic potash, a yellow precipitate falls which has the composition of red oxide of mercury, the reaction being as follows:—

$$\underset{\text{Mercuric chloride.}}{HgCl_2} + \underset{\text{Caustic potash.}}{2KOH} = \underset{\text{Mercuric oxide.}}{HgO} + \underset{\text{Potassium chloride.}}{2KCl} + \underset{\text{Water.}}{H_2O}.$$

This compound is a very finely divided precipitate, but otherwise is identical with the red oxide. It is distinguished in the pharmacopœia under the name "*hydrargyri oxidum flavum*, or yellow oxide of mercury."

Red (and yellow) oxide of mercury is poisonous, the antidotes being, albumen, mucilaginous drinks, etc., as with corrosive sublimate.

Mercuric Nitrate.—Formula, $Hg(NO_3)_2$. Molecular weight, 323.58.

Synonym.—Nitrate of mercury.

This salt is made by dissolving mercury in hot nitric acid. It is of no particular interest to the pharmacist or physician, as it is not used in medicine, if we may except its problematical existence in citrine ointment. If it be desired pure, mercury must be boiled in nitric acid until a portion of the solution ceases to produce a precipitate when dropped into solution of sodium chloride. The solution stains the skin a dark red. When strong nitric acid is added to a concentrated solution of mercuric nitrate, crystals (R. & S.) of the composition $Hg_2(NO_3)_2 + H_2O$, separate.

Mercuric Sulphate.—Formula, $HgSO_4$. Molecular weight, 295.62.

Synonym.—Sulphate of mercury.

This salt is made by heating mercury in sulphuric acid.

Experiment.—Heat one part of mercury in a porcelain dish on a sand-bath, with one and one-half parts of sulphuric acid, stirring constantly with a porcelain spatula, until the mercury disappears and a white powder remains. It is necessary that this reaction be performed in the open air or beneath a good flue, that the sulphur dioxide (sulphurous acid gas) evolved may escape, together with excess of sulphuric acid. The reaction is as follows:—

$$\underset{\text{Mercury.}}{Hg_2} + \underset{\text{Sulphuric acid.}}{4H_2SO_4} = \underset{\text{Mercuric sulphate.}}{2HgSO_4} + \underset{\text{Water.}}{4H_2O} + \underset{\text{Sulphur dioxide.}}{2SO_2}.$$

Sulphate of mercury is a white powder, soluble in sulphuric acid. It turns yellow upon heating. Sulphate of mercury is seldom, if ever, used in medicine, but of late years has come into demand with physicians, as it is employed with some galvanic batteries. When boiled with water it decomposes, a yellow basic salt (oxysulphate) remaining, Hg_3SO_6. This is officinal in the pharmacopœia as

Hydrargyri Sulphas Flava.

Synonyms.—Yellow sulphate of mercury. Turpeth mineral.

It is made by boiling four parts of mercury with six parts of sulphuric acid and then pouring the white powder (sulphate of mercury) into boiling water, washing the precipitate repeatedly with water, and drying it. Turpeth mineral is but little used at the present day.

Mercuric Sulphide.—Formula, HgS. Molecular weight, 231.78.

Synonyms.—Hydrargyri sulphuretum rubrum. Red sulphuret of mercury. Red sulphide of mercury. Cinnabar. Vermillion.

Sulphide of mercury occurs native, often as distinct crystals, and is then known under the name *cinnabar*. It is made artificially, both by wet processes and by sublimation. To make it by sublimation, mercury is triturated with flowers of sulphur in proper proportion, until a brown powder results, then the mixture is sublimed. By the wet process, the brown powder obtained by rubbing mercury with sulphur is boiled with a solution of caustic potash until the precipitate reaches the proper color, when it is at once washed in cold water and dried. Manufacturers use various proportions of sulphur and mercury. The British Pharmacopœia recognizes forty parts of mercury and eight parts of sulphur, mercury being poured into the melted sulphur. Many different shades of vermillion are made, and they depend in part upon the proportion of free sulphur present. Sulphide of mercury dissolves in concentrated hydriodic acid, sulphide of hydrogen being liberated. It also dissolves in nitro-muriatic acid, with

liberation of sulphur. If it sublimes perfectly the usual adulterants are absent, for red lead and oxide of iron will remain. Some chemists admit the existence of—

Mercurous Sulphide, Hg_2S, but exception is taken to this by excellent authorities, who protest against the existence of such a compound, stating that the black sulphide is simply an amorphous form of the red.†

Black sulphide of mercury is made by triturating together equal weights of sulphur and mercury, until the globules of mercury disappear, and a black heavy powder remains. This is known as *Æthiops mineral* and may well be accepted as a mixture of sulphur and amorphous *mercuric sulphide.*

Tests for mercury.—Place a strip of clean zinc partly covered with gold foil in a solution of any mercury salt. The gold will become covered with a thin white metallic film. Heat the strip to redness, and the mercury is volatilized.

Acidulate the solution of a salt of mercury with a little hydrochloric acid and immerse in it a strip of bright copper. Mercury will be deposited as a silver-like coating, and upon heating the copper to redness the film will disappear.

Sulphide of hydrogen forms precipitates with solutions of any salt of mercury. This is insoluble in excess of the precipitant and is black. It is insoluble in sulphide of ammonium. To distinguish this from other black sulphides, boil the precipitate with dilute nitric acid and then wash it well with water and dissolve it in the minimum amount of nitro-muriatic acid, and add a slight excess of ammonia water, and then a slight excess of hydrochloric acid. Examine this solution with the metallic copper test.

Mercurous salts form white precipitates with soluble chlorides, and these turn black upon the addition of ammonia. They precipitate black with caustic potash or caustic soda.

† "*Mercurous sulphide* does not exist, a mixture of metallic mercury and mercuric sulphide being produced in cases in which its formation might have been expected." (Roscoe & Schorlemmer.)

Mercuric salts precipitate yellow with caustic potash or caustic soda. They yield scarlet precipitates if carefully added to solution of potassium iodide, but this precipitate dissolves in excess of either reagent.

SILVER (*Argentum*).

Symbol, Ag. Atomic weight, 107.66.

Silver occurs native, sometimes in large masses. It is usually present in lead and copper ore in small proportion. Usually native silver is contaminated with gold and copper and other metals. The waters of the ocean also contain silver. It is unnecessary for us to notice the processes employed for the extraction of silver.

Silver (bullion) of commerce is never pure, always containing traces of other metals. Silver is the best known conductor of electricity. It is soft, pure white, tough and malleable, can be highly polished, and crystallizes in form of the regular octohedron. When heated to whiteness in the oxyhydrogen flame, silver volatilizes, boils, and completely distills. Silver dissolves in nitric acid, but not in hydrochloric acid. Silver tarnishes upon exposure to the air, consequent upon the formation of sulphide of silver. For this reason silver ware turns black in the neighborhood of sulphur springs, as do the coins upon the persons of those exposed to sulphide of hydrogen. The specific gravity of the various forms of silver are given by investigators as follows:—

Distilled Silver	10.575
Fused Silver	10.424 to 10.511
Pressed Silver	10.57

Silver is employed extensively in the arts for electroplating etc. Silver coin is an alloy containing various proportions of copper, according to the country whence it is issued. Metallic silver is not used in medicine, but is employed for silvering surgical instruments and as a fine wire or thread for surgeons use. Few of the salts of silver are employed in medicine.

Silver Cyanide.—Formula, AgCN. Molecular weight, 133.64.

Synonyms.—Argenti cyanidum. Cyanide of silver.

This salt is officinal in the U. S. P., but is never administered internally as a remedy. It is made according to the above authority by a circuitous process, hydrocyanic acid being first prepared and this decomposed with nitrate of silver. It may be made by dissolving a little nitrate of silver in distilled water, and then adding dilute hydrocyanic acid to the solution until in excess A white curdy precipitate at once forms which must be washed, collected on a filter paper, and dried (all in the dark) by exposure to the atmosphere. The reaction is as follows:—

$$\underset{\text{Silver nitrate.}}{AgNO_3} + \underset{\text{Hydrocyanic acid.}}{HCN} = \underset{\text{Silver cyanide.}}{AgCN} + \underset{\text{Nitric acid.}}{HNO_3}.$$

Cyanide of silver is white, but turns brown by exposure to light. It dissolves in ammonia water, but not in dilute nitric acid. It forms soluble double salts with cyanides of the alkali metals and with cyanides of the alkaline earth metals. Hydrochloric acid decomposes it, chloride of silver being formed and hydrocyanic acid evolved, thus:—

$$\underset{\text{Silver cyanide.}}{AgCN} + \underset{\text{Hydrochloric acid.}}{HCl} = \underset{\text{Silver chloride.}}{AgCl} + \underset{\text{Hydrocyanic acid.}}{HCN}.$$

If cyanide of silver be mixed with dilute sulphuric acid, and strips of metallic zinc be placed in the mixture, the zinc will dissolve, the cyanide of silver decompose, and metallic silver will precipitate. This is generally considered pure silver when washed and dried, but it is almost an impossibility to separate the adhering substances by this method. Stas has suggested the employment of caustic potash and milk-sugar for the purpose of reducing the silver from a very pure cyanide. Cyanide of silver is only employed in a pharmaceutical sense for the extemporaneous preparation of dilute hydrocyanic acid.

Silver Nitrate.—Formula, $AgNO_3$. Molecular weight, 169.55.

Synonyms.—Argenti nitras. Nitrate of silver.

This has been known from a very early day, and is used extensively in the arts, in photography, and in medicine. Nitrate of silver may be easily made by dissolving pure metallic silver in dilute nitric acid, evaporating the solution to crystallization, and afterward dissolving the crystals in water and recrystallizing. According to the U. S. P., two parts of silver are dissolved in two and one-half parts of nitric acid previously mixed with one part of water, heat being employed. The solution is then evaporated to dryness, heat being applied until the mass melts, when it is stirred until nitric acid vapors cease to escape. It is then cooled, dissolved in water, the solution filtered, and the filtrate evaporated to dryness. The object of the heat in the above case is to rid the salt of copper.

Nitrate of silver crystallizes in transparent plates of the rhombic system. The specific gravity is 4.328 and its melting point 198° C. (388.4° F.) It is permanent in the air when pure, unless organic matter is present, when it blackens. The taste is acrid and metallic, and in large doses it acts as a corrosive poison, destroying the membranes with which it comes in contact. The crystals are odorless, soluble in their weight of cold water, and are more soluble in hot water. When ammonia water is added to a solution of nitrate of silver, a precipitate first forms, then, as the ammonia is increased, it redissolves. This solution deposits crystals of the composition, $AgNO_3 + 2NH_3$, upon evaporation. The foregoing ammoniated solution of nitrate of silver is often used in connection with a solution of pyrogallic acid, to color hair black. Nitrate of silver is a strong poison, and the antidote should be a soluble chloride, sal-ammoniac, or common salt, brine being usually convenient. The stomach pump, or emetics should be also employed, and followed by mucilaginous drinks.

Silver Oxide. — Formula, Ag_2O. Molecular weight, 231.28.

Synonyms.—Argenti oxidum. [Silver hemioxide.]

This is easily prepared by the simple process of the U. S. P., as follows:—

"Take of nitrate of silver, four troy ounces; distilled water half a pint; solution of potassæ, a pint and a half, or a sufficient quantity. Dissolve the nitrate of silver in the water, and to the solution add solution of potassæ as long as it produces a precipitate. Wash this with water until the washings are nearly tasteless. Lastly, dry the precipitate, and keep it in a well stoppered bottle, protected from the light."

The above is explained by the following equation:—

$$\underset{\text{Silver nitrate.}}{2AgNO_3} + \underset{\text{Potassium hydrate.}}{2KOH} = \underset{\text{Silver oxide.}}{Ag_2O} + \underset{\text{Potassium nitrate.}}{2KNO_3} + \underset{\text{Water.}}{H_2O}.$$

Oxide of silver is a brown or almost black powder. It dissolves in 3000 parts of water, rendering the solution alkaline, reacts with acids to form salts, decomposes partly at 250° C., completely at 300° C., parting with all its oxygen, pure silver remaining. It should never be made into pills with essential oils, tannin, and like organic substances, nor rubbed with sulphur or the metallic sulphides or phosphides; as with these or other easily oxidizable substances combustion may ensue. The substance known as—

Fulminating silver, or ammonia-silver oxide, is made by dissolving silver oxide in ammonia water. The actual composition of this substance has never been ascertained, few caring to experiment even with very small amounts of it. Fulminating silver is among the most dangerous of explosive substances, a slight jar sufficing to decompose it, the detonation being terrible from even a few grains. It should *never* be made by the experimenter.

Silver tetrantoxide (argentous oxide), Ag_4O_3, is made from silver citrate by heating it in a current of hydrogen, and,

Silver peroxide or dioxide, Ag_2O_2, produced by the action of ozone on silver oxide (Ag_2O), or when a galvanic

current is passed through a solution of nitrate of silver. Neither is of interest.

Many other salts of silver are known, few, if any, of interest to the pharmacist. We may briefly mention the following:—

Silver Bromide, AgBr, found native in Chili and Mexico. This salt may be made by precipitating a solution of nitrate of silver with a solution of hydrobromic acid and drying the precipitate. The precipitate is white, but darkens upon exposure to light, therefore it should be prepared in darkness Bromide of silver dissolves in hot hydrobromic acid, concentrated ammonia, and solution of nitrate of mercury. It melts by heat, turning red.

Silver Chloride, AgCl. This is the mineral known under the name *horn-silver*, occurring in many parts of the world. It may be easily made by adding hydrochloric acid to solution of nitrate of silver, the chloride of silver separating as a white curd. It must be dried in darkness, for exposure to light turns it gray.

Chloride of silver is white, insoluble in water, but freely soluble in ammonia water (forming a double salt, *ammonio-chloride of silver*, $2AgCl+3NH_3$), and in solution of chloride of ammonium, and in solution of chloride of sodium. When mixed with dilute hydrochloric or sulphuric acid, metals, such as iron or zinc, if immersed, will precipitate metallic silver upon their surface. The specific gravity of chloride of silver is 5.5.

Silver Iodide, AgI, occurs as a mineral in many localities, but may be made by direct combination of silver and iodine, or by dissolving metallic silver in hydriodic acid, thus:—

$$?HI + 2Ag = 2AgI + 2H.$$

Hydriodic acid. Silver. Silver iodide. Hydrogen.

Iodide of silver may best be made by double decomposition between solution of iodide of potassium and solution of nitrate of silver, the result being iodide of silver (which precipitates yellow) and nitrate of potassium thus:—

$$KI + AgNO_3 = AgI + KNO_3$$

Potassium iodide. Silver nitrate. Silver iodide. Potassium nitrate.

Iodide of silver, when dried, is a bright yellow powder, of specific gravity 5.687 (precipitated). It contracts by heating until the temperature of 70° C. (158° F.) is reached, and expands again on cooling. It is not altered, if pure, by light. Ammonia dissolves it sparingly.

Silver Phosphate, Ag_3PO_4. This is the *normal* phosphate of silver, and is made by pouring a solution of phosphate of sodium into a solution of nitrate of silver until precipitation ceases. It is yellow, and insoluble in water. This salt is of no importance.

Pyrophosphate of Silver, $Ag_4P_2O_7$, is white, and may be made by double decomposition between solutions of pyrophosphate of sodium and nitrate of silver.

Silver Sulphate, Ag_2SO_4. This salt may be made by digesting carbonate of silver with dilute sulphuric acid, but is most readily made by double decomposition between concentrated solution of nitrate of silver and of sulphate of sodium, the precipitate being washed with water and dried. The reaction is expressed by the following equation:—

$$\underset{\text{Sodium sulphate.}}{Na_2SO_4} + \underset{\text{Silver nitrate.}}{2AgNO_3} = \underset{\text{Silver sulphate}}{Ag_2SO_4} + \underset{\text{Sodium nitrate.}}{2NaNO_3}.$$

Sulphate of silver is a white, odorless, crystalline powder, soluble in 200 parts of cold water, and about 70 parts of hot water. It dissolves freely in dilute sulphuric or nitric acid. Light gradually changes it to gray with decomposition.

Silver Sulphide, Ag_2S, may be made by passing sulphide of hydrogen into a solution of nitrate of silver. It is black, insoluble in water, but soluble in hot nitric acid. The brown or yellowish stain on silver ware is usually caused by the formation of this compound, and coins in the pockets of workmen exposed to sulphide of hydrogen often become black. Silver spoons blacken from the sulphur of eggs. Tea will also blacken silver ware, the result in this case being *tannate* of silver. Ammonia will remove this latter stain.

The soluble salts of silver are poisonous, the antidote being in all cases that recommended for nitrate of silver, (see page 280.)

Tests for Silver.—To the solution of a salt of silver, add hydrochloric acid until it ceases to produce a precipitate. Boil this precipitate with an excess of distilled water (to dissolve lead chloride), and collect the insoluble precipitate on a filter. Add this precipitate to ammonia water. It will dissolve, if it be pure silver chloride. If a gray mass remain (mercurous chloride) filter from it the ammoniacal solution, and to the filtrate add, very carefully, nitric acid until the excess of ammonia is neutralized; there will then occur a white precipitate of silver chloride. If the silver salt be insoluble, mix it with sodium carbonate and fuse before the blow-pipe flame, on a charcoal support. A bead of silver results which can be dissolved in nitric acid, and this solution examined as before stated.

CERIUM GROUP OF METALS.

CERIUM,	LANTHANUM,
DIDYMIUM,	TERBIUM,
ERBIUM,	YTTRIUM.

The elements of this group are rare, and excepting cerium, of no value in the arts or medicine. It is very troublesome to separate them from each other. They are derived from minerals, mostly rare, obtained from Greenland, Scandinavia, and Siberia. Brief reference shall be made of each of the elements, but it is unnessary in this work to consume space with the processes adapted by experimenters for their extraction, or to notice their salts.

CERIUM.

Symbol Ce. Atomic weight, 141.2.

Cerium was discovered by Klaproth (1803) and simultaneously by Berzelius and Hisinger, the source being a Swedish mineral. Mosander obtained it by heating chloride of cerium with sodium, and afterward Hillebrand and Norton by electrolysis of the chloride. In

color cerium resembles iron, it tarnishes in moist air, passing through various shades of yellow, blue and green. It is hard, malleable when warm, and may be drawn into wire. It melts between the melting points of silver and antimony. It burns in the air like magnesium, but more brilliantly, and when the metal is struck by flint the particles separated scintillate beautifully. The specific gravity of cerium is 6.628 before fusing, 6.728 after melting. Cerium forms two oxides:—

Cerium Sesquioxide (cerous oxide), Ce_2O_3, a bluish green powder, and—

Cerium Dioxide (ceric oxide), CeO_2, white or straw colored. Neither is of interest. Cerium forms two series of salts called *cerous salts* and *ceric salts*; they correspond with the oxides. The cerous salts are best understood. *Cerous oxalate* is the only salt of this element that is employed as a medicine.

Cerium Oxalate.

Synonyms.—Oxalate of cerium. Cerii oxalas.

Oxalate of cerium is made by a tedious process from the mineral *cerite*. For its production the powdered mineral is digested with hydrochloric acid. The solution is evaporated to small bulk, treated with sulphide of hydrogen, diluted and filtered. Ammonia water is then added in slight excess, the precipitate collected, washed with water, dissolved by means of dilute hydrochloric acid, and precipitated by solution of oxalic acid. This produces crude oxalate of cerium, from which the impurities (lanthanum and didymium) are separated by calcining it, whereby sesquioxide of cerium is produced with the oxides of the other metals. The powder is then digested with solution of sal-ammoniac, which dissolves all but the cerous oxide. This is then dissolved in dilute hydrochloric acid, and oxalic acid solution added until in slight excess. The precipitate of oxalate of cerium is then dried. It will be observed that the object of the foregoing roundabout process is to obtain *chloride* of *cerium*, and that from this the oxalate is in reality prepared by decomposition with oxalic acid.

Oxalate of cerium is a white powder, soluble in sulphuric and hydrochloric acids. It decomposes when heated in a current of hydrogen gas, *cerous* oxide (sesquioxide) remaining. It is tasteless, but grits between the teeth, seeming to make them adhere together when moved upon each other. It should not effervesce when dropped into dilute acids. The other salts of cerium do not require attention from us.

DIDYMIUM.

Symbol, Di (D). Atomic weight, 147.0.

This metal was prepared by Marignac in the manner in which cerium was produced by Morlander, and afterward was obtained by Hillebrand and Norton, from the fused chloride by electrolysis. Didymium resembles cerium in appearance, but is darker, and inclined to a yellowish tint. It burns in the air; has the specific gravity of 6.544; forms an oxide of the composition, Di_2O_3, and gives rise to a series of sweetish, rose colored salts, of no general interest or importance.

ERBIUM.

Symbol, Er. Atomic weight, 169.0.

This element has never been isolated, but its oxide (Er_2O_3), under the name *erbia*, is familiar to experimental chemists. A series of rose colored salts are known; they react acid, and are astringent to the taste.

LANTHANUM.

Symbol, La (L). Atomic weight, 139.0.

Lanthanum was prepared by heating the chloride of this element with potassium (Mosander), and by electrolysis of the fused chloride (Hillebrand and Norton). It is an iron-gray metal, tarnishes in the air, is malleable, and burns when heated in a flame. It combines with both chlorine gas and bromine vapor, evolving heat and light, oxidizes in water and dissolves quickly in diluted sulphuric and diluted hydrochloric acids, hydrogen gas being rapidly

evolved. It is not affected by concentrated oil of vitriol. The salts of lanthanum are colorless and have never been applied to any practical use.

YTTRIUM.

Symbol, Y. Atomic weight, 93.0

This metal (mixed with erbium) was obtained by Berzelius by fusing potassium with the chloride. It yields an oxide, Y_2O_3, and a line of salts which are of no value.

ALUMINIUM GROUP OF METALS.

ALUMINIUM. GALLIUM. INDIUM.

These elements form sulphates, and these sulphates unite with sulphates of the alkali metals to produce double salts, which crystallize in regular octohedrons. Aluminium salts are of value in medicine and the arts, but the salts of gallium and indium are of no importance.

ALUMINIUM.

Symbol, Al. Atomic weight, 27.3.

Alum has been known from an early period, but until 1827 the element which entered into its composition was unknown; then Wœhler prepared it by fusing chloride of aluminium with potassium. In combination, it is among the most abundant of elements, being found in almost all rocks and soils, and composing the bulk of many minerals or even mountains. Clays are disintegrated felspars, colored with metallic oxides. Strange as it may seem, the higher orders of plants do not contain even traces of any compounds of aluminium, but certain cryptogamic plants yield it in abundance. Aluminium is now made in considerable amounts by mixing together alumina, chloride of sodium, and coal dust; heating the mixture in upright furnaces and passing into it chlorine gas. Double

chloride of aluminium and sodium distil, and after condensation is mixed with cryolite (aluminium and sodium fluoride) and sodium, and the mixture is fused, when aluminium more or less impure separates in the molten state. Aluminium is white and may be polished. When pure it does not oxidize in the air, it conducts electricity better than iron, fuses at about 700° C. and dissolves in solutions of the caustic alkalies, hydrogen escaping. It dissolves freely in hydrochloric acid, slightly in dilute sulphuric, and readily in solutions of the cyanic acids mixed with solution of salt. Aluminium forms an alloy with copper which can scarcely be distinguished from gold, and which is used in making philosophical apparatus and jewelry. Aluminium is very light, having the specific gravity of but 2.67, and for this reason is used for making small weights. The common salt of aluminium is—

Alum. (Ammonium alum).—Formula. $(NH_4)_2Al_24SO_4.24H_2O$. Molecular weight. 904 94.

Synonyms. — Aluminium and ammonium sulphate. Aluminii et ammonii sulphas. Sulphate of aluminium and ammonium. Ammonia-alum. Alum. Alumen.

Ammonia alum is now officinal in the U. S. Pharm. and is made in practice by digesting clay with sulphuric acid, lixivating the mass with water, filtering, and adding sulphate of ammonium† to the filtrate when crystals of ammonia-alum form. The properties of this are almost identical with those of potash-alum, but when it is triturated with a caustic alkali the odor of ammonia is evolved, and when it is heated to redness alumina remains, with escape of ammonia. Both this, and potash alum, are usually contaminated with iron, which strikes a blue color with ferrocyanide of potassium, but the impurity is generally unimportant. Ammonia-alum dissolves in about eight parts of water at 20° C (68° F.) and rather less than one-fourth part of water, at the boiling point of water.

† Sulphate of ammonium is now abundantly produced as a side product of gas manufacture, which accounts for the cheapness of the ammonia-alum.

Aluminium and Potassium Sulphate. (Potash alum.)
Formula, $K_2Al_24SO_4.24H_2O$. Molecular weight, 947.

Synonyms.—Aluminii et potassii sulphas. Potassa alum. Sulphate of aluminium and potassium. Alum. Alumen.

Potash alum was officinal in the Pharmacopœia of the United States, 1860, but the alum employed in medicine and the alum of commerce at this day, is the ammonium alum, now officinal. The composition of potash alum is represented at the beginning of this article, but some chemists prefer to express it, $Al_2(SO_4)_3 + K_2SO_4 + 24H_2O$. Potash alum is made from alum-stone (the mineral alunite), and others of like nature, and from clay, by the action of sulphuric acid, which forms sulphate of aluminium. This is dissolved in water, and subsequently mixed with sulphate of potassium, which produces the double salt. For many years ammonia-alum was also principally used in Europe, but the potash beds of Stassfurt supply chloride of potassium so abundantly that potash-alum has superseded ammonia-alum, even in England. Potash-alum crystallizes in transparent regular octohedrons. It is astringent to the taste, insoluble in alcohol, but is very soluble in boiling water (one part of water dissolving about three and one-half parts of alum). One part of cold water dissolves rather more than one-fourth part of alum. The crystals absorb ammonia and turn white on the surface when exposed to the air, melt at 92° C. (197.6° F.), and at above 100° C (212° F.) part with all their water of crystallization, "burnt alum" remaining. When it is heated to redness the odor of ammonia is not apparent, and the residuum consists of a mixture of sulphates of aluminium and potassium. Potash-alum does not evolve the odor of ammonia when it is triturated with caustic potash. When an alkali is slowly added to solution of alum the precipitate thrown down redissolves to a certain point; when the precipitate ceases to redissolve, if the addition of alkali be suspended it will be found that the solution is neutral. This is used in the arts for dyeing, under the name *neutral alum.*

Dried Alum.

Synonyms.—Alumen exsiccatum. Burnt alum.

"Take of alum, in coarse powder, four troy ounces, place it in a suitable vessel, and subject it to a temperature not exceeding 400°, until the residue weighs two troy ounces and one hundred and twenty grains; then, when cold, reduce it to a fine powder."—*U. S. P.*

By this process the largest part of the water of crystallization is expelled. Burnt alum is a white, porous, light substance, very astringent, and is considerably employed in medicine. When burnt alum is mixed with bicarbonate of sodium, and the mixture is moistened, carbon dioxide escapes and sulphates of both sodium and aluminium are formed. Advantage is taken by some of this fact to prepare cheap baking-powder. The use of this, however, should be discountenanced, inasmuch as so much aluminium sulphate is objectionable in food.

Aluminium Sulphate.—Formula, $Al_2(SO_4)_3$ Molecular weight, 342.06. Crystallized — $Al_2(SO_4)_3.18H_2O$. Molecular weight, 665.34.

Synonyms.—Aluminii sulphas. Sulphate of aluminium. Concentrated alum.

This preparation is officinal in the pharmacopœia and is prepared by mixing solutions of equal weights of alum and carbonate of sodium, washing the precipitate well with water and then dissolving it by means of sulphuric acid and evaporating this solution to crystallization. The process is explained as follows; alum, and sodium carbonate when mixed yield hydroxide of aluminium, thus:—

$$\underset{\text{Alum.}}{Al_2(NH_4)_2(SO_4)_4} + \underset{\text{Sodium carbonate.}}{3Na_2CO_3} + \underset{\text{Water.}}{3H_2O} =$$

$$\underset{\substack{\text{Aluminium}\\\text{hydroxide.}}}{Al_2(OH)_6} + \underset{\substack{\text{Sodium}\\\text{sulphate.}}}{3Na_2SO_4} + \underset{\substack{\text{Ammonium}\\\text{sulphate.}}}{(NH_4)_2SO_4} + \underset{\substack{\text{Carbon}\\\text{dioxide.}}}{3CO_2}.$$

The sulphates of sodium and ammonium are washed from the precipitated aluminium hydroxide, which is then dissolved by means of sulphuric acid, forming sulphate of aluminium in the following manner:—

$$\underset{\text{Aluminium hydroxide.}}{Al_2(OH)_6} + \underset{\text{Sulphuric acid.}}{3H_2SO_4} = \underset{\text{Aluminium sulphate.}}{Al_2(SO_4)_3} + \underset{\text{Water.}}{6H_2O.}$$

Sulphate of aluminium is a white powder, soluble in its weight of water and of a peculiar acidulous, astringent, not unpleasant taste. The solution precipitates with a small amount of caustic potash, but this precipitate redissolves in an excess of that article. It melts when heated in its water of crystallization.

Aluminium Hydroxide, or *hydrate of aluminium,* $Al_2(OH)_6$, is of interest, since it may be employed to prepare other salts of aluminium in the manner given for making sulphate of aluminium. Thus, if the gelatinous precipitate, $Al_2(OH)_6+2H_2O$, be dissolved in cold acetic acid,

Acetate of Aluminium results, as explained by the following equation:—

$$\underset{\text{Aluminium hydroxide.}}{Al_2(OH)_6} + \underset{\text{Acetic acid.}}{6HC_2H_3O_2} = \underset{\text{Aluminium acetate.}}{Al_26C_2H_3O_2} + \underset{\text{Water.}}{6H_2O.}$$

Acetate of aluminium is also made (impure) by double decomposition of acetate of lead and alum. It is used as a mordant in dyeing cloth.

Chloride of Aluminium, Al_2Cl_6, results when hydroxide of aluminium is dissolved in dilute hydrochloric acid, the reaction being explained as follows:—

$$\underset{\text{Aluminium hydroxide.}}{Al_2(OH)_6} + \underset{\text{Hydrochloric acid.}}{6HCl} = \underset{\text{Aluminium chloride.}}{Al_2Cl_6} + \underset{\text{Water.}}{6H_2O.}$$

This salt is hygroscopic and has the composition, when crystallized, of $Al_2Cl_6.12H_2O$. In an impure form it is used in solution as a disinfectant and deodorizer. There are many other compounds of aluminium which are employed in the arts, chiefly as mordants, but none are of interest to the physician.

GALLIUM.

Symbol, Ga. Atomic weight, 69.8.

This metal is a late discovery, having been noticed for the first time by L. de Boisbaudran, in the year 1875.

He detected it by means of spectrum analysis, and its separation was effected by means of electrolysis of the alkaline solution.

Metallic gallium has a bluish-white color, is tough, oxidizes upon exposure to the air, and has the specific gravity of 5.9. It does not volatilize at redness. Gallium melts at 30.1° C. (about 86° F.), and the molten metal refuses to harden even when reduced to 0° C. (32° F.) unless it be touched with a piece of solid gallium, when it immediately solidifies. Gallium forms an oxide, Ga_2O_3. A few of its salts have been studied, but as yet the compounds of the element are little known, and of no use.

INDIUM.

Symbol, In. Atomic weight, 113.4.

This metal was discovered in 1863 by Reich and Richter. It was detected by spectrum analysis, the origin being certain zinc-blendes. The method of production is tedious and unnecessary for us to notice. Indium is white, malleable, softer than lead, and of the specific gravity 7.42. It melts at 176° C. (348.8° F.), does not oxidize in the air, and forms an oxide, In_2O_3, and a hydroxide, $In_2(OH)_6$. There is a series of indium salts, but they are of interest neither in the arts nor medicine.

THE METALS OF THE IRON GROUP.

COBALT.	MANGANESE.
IRON.	NICKEL.

Three of the group are employed in medicine, cobalt being the exception.

COBALT.

Symbol, Co. Atomic weight, 58.6.

Cobalt is not found pure. The ores of cobalt have been employed for a considerable period to color glass blue. In 1735 Brandt identified the element, and Bergman and others afterward supported his discovery.

The ores of cobalt are found in a limited number of places and are very impure, containing large quantities of arsenic, iron, nickel, and other elements. The chief ore, an impure arsenide of cobalt, is sold in commerce under the name, *cobalt* or *fly-stone*, for the purpose of poisoning flies. Metallic cobalt is of no importance. It resembles polished iron; has a specific gravity of from 8.5 to 8.7; is tough and malleable, and harder than iron.

Cobalt forms two oxides—

Cobalt monoxide.....CoO.
Cobalt sesquioxide.....Co_2O_3.

Neither is employed in medicine. The former (cobalt monoxide) gives rise to a series of salts known as *cobaltous salts*. The latter (cobalt sesquioxide) forms *cobaltic salts*. The ordinary salts of cobalt are of the first series, and as such we may name chloride of cobalt, $CoCl_2$, nitrate of cobalt, $Co(NO_3)_2.6H_2O$, etc. Cobalt salts are employed in preparing the so-called sympathetic inks, for when free from moisture they are usually of a blue color, but turn pink when moist. When chloride of cobalt is made into dilute solution the writing is invisible until warmed, when it becomes blue, the color disappearing when the paper is cooled.

Cobalt forms a number of ammoniacal compounds known as *cobaltamine salts*, which are all characterized by peculiar colors, red, blue, green, pink, and intermediate shades. Cobalt unites with sulphur, with phosphorus and with arsenic to form their respective compounds.

IRON. (*Ferrum.*)

Symbol, Fe. Atomic weight, 55.9.

This element is seldom found in the metallic form, but its ores are very extensively distributed and in great abundance. It is almost always present in meteoric stones. Plants contain, and seem to demand it, and animals find a small portion of iron necessary to existence; the red coloring matter of blood contains one-fourth of one per cent.

Iron of commerce is usually very impure, and contains, notably, carbon in greater or less proportion. Pure iron is produced by electrolysis; by reduction of *ferrous chloride* in hydrogen gas, and by other processes which are tedious and unnecessary to name.

Pure iron is almost silver-white, and is easily polished. It has the specific gravity 7.84, and is very tenacious. It softens at redness, and at a white heat can be melted. It volatilizes in a vacuum under the influence of a strong electrical discharge, and this is, to an extent, the case when it is burned in oxygen gas. Iron is attracted by the magnet, and in turn may be made magnetic. Neither dry oxygen, air, nor pure water free from air, affects it; but moist air, or water containing air, soon covers it with a coating of rust (ferric hydroxide). Iron burns in oxygen gas when heated, also in vapor of sulphur, and combines at high temperatures with carbon. All commercial iron contains carbon. Cast-iron contains the largest amount, wrought-iron the smallest, and steel proportions between these extremes. Iron precipitates copper from solution, dissolves in many dilute acids, and forms two basic oxides which in turn unite to produce the well known magnetic oxide of iron (ferroso-ferric oxide). The salts of iron are very numerous, and in medicine are, undoubtedly, employed in larger amount than the salts of any other metal, and in addition are invaluable in the arts and manufactories. A form of metallic iron in very finely divided state is used in medicine and known as—

Reduced Iron.

Synonyms.—Ferrum redactum. Iron reduced by hydrogen. Iron by hydrogen. Quevenne's iron.

This is made according to the U. S. P. essentially as follows:—

Officinal sub-carbonate of iron is thoroughly washed with water and then heated until it is free from moisture. This is then placed in thin layer on a tray and the tray is placed in an iron cylinder. Hydrogen gas is then made, purified, and passed into the cylinder, and this is then heated, by means of a charcoal furnace, to dull red-

ness, and the flow of hydrogen gas continued until the iron is completely reduced.

The foregoing is explained from the fact that the heated hydrogen deprives the ferric oxide (such in reality being the contents of the tube) of its oxygen, forming water which escapes, the iron being left as a fine powder, thus:—

$$\underset{\text{Ferric oxide.}}{Fe_2O_3} + \underset{\text{Hydrogen.}}{H_6} = \underset{\text{Iron.}}{Fe_2} + \underset{\text{Water.}}{3H_2O}.$$

Iron by hydrogen is "a very fine gray powder without gloss; when heated in the air it burns to ferric oxide. It is completely soluble in warm, diluted hydrochloric acid, with the evolution of hydrogen, which is indifferent to lead paper. When treated for half an hour, at the ordinary temperature and with occasional agitation, with twenty-five times its weight of solution of ferric chloride, specific gravity 1.30, it is completely dissolved." (Schacht.)

According to the U. S. P., it should be readily ignited with a lighted taper. Iron by hydrogen usually contains magnetic oxide of iron (Fe_3O_4), which imparts a black color to the powder, and both the German and the British Pharmacopœias permit the presence of fifty per cent. of this article.

Iron forms two classes of salts which are known respectively as *ferric* salts and *ferrous* salts. The ferric salts contain the least proportion of iron, thus:—Fe_2Cl_6 is *ferric* chloride, while $FeCl_2$ is *ferrous* chloride. By comparing the foregoing it will be seen that *ferric* chloride contains two atoms of iron united with six atoms of chlorine, while the *ferrous* chloride contains one atom of iron in combination with two atoms of chlorine. When iron is simply acted upon by dilute acids *ferrous* salts are usually produced. From this reason we shall first consider these *lower* salts of iron, as the ferrous class is often designated.

Iron wire is generally directed as the form to be used in pharmaceutical and chemical operations.; but iron turnings from the iron foundry may be used instead. In such cases care must be taken that brass or copper is not present, and it is *always* necessary to wash the turn-

ings well in alkali water, to purify them from the grease which is employed in drilling. Wrought-iron is preferable to either cast-iron or steel.

FERROUS SALTS.

Ferrous Arseniate.—Formula, $Fe_3 2AsO_4$. Molecular weight, 445.18.

Synonyms.—Ferri arsenias. Arseniate of iron.

This compound is but little employed in this country, however, it is officinal in the British Pharmacopœia.

It is made by dissolving three parts of acetate of sodium and four parts of dried arseniate of sodium in thirty-two parts of boiling distilled water, and mixing this solution with a hot solution of nine parts of ferrous sulphate in forty-eight parts of distilled water, and drying the precipitate after having washed it well.

When the solutions are mixed a white precipitate is produced of the composition expressed by our formula, $Fe_3 2AsO_4$. This is not permanent, however, and as it dries, absorbs oxygen from the air, passes through various shades of green and blue, finally becoming of a blue-green, very nearly resembling the dry precipitated phosphate of iron. It now has the composition according to Maisch, $2Fe_3As_2O_8.4FeAsO_4\ Fe_2O_3.32H_2O.$, and is a ferroso-ferric arseniate. Arseniate of iron dissolves in many dilute acids. It is tasteless and odorless, and the solution in acids yields blue precipitates with both the red and the yellow prussiate of potash. It may be mistaken for the blue phosphate of iron, which it nearly resembles. In case it is necessary to apply a test, boil a small portion of the powder with an excess of caustic soda solution, filter, neutralize the filtrate exactly with nitric acid, and then add solution of nitrate of silver. If *arsenic* is present a brick-red precipitate will be produced. This salt is poisonous and cases of poisoning should be treated according to the directions we have given under arsenious acid.

Ferrous Bromide.—Formula, $FeBr_2$. Molecular weight, 215.4.

Synonyms.—Bromide of iron. Ferri bromidum.

This may be made by direct combination between bromine and iron, or by dissolving metallic iron or red oxide of iron in hydrobromic acid. The following from the supplement to the American Dispensatory is practical.

"Into a chemical flask introduce two fluid ounces of distilled water and one troy ounce of clean coarse iron turnings, or iron wire, and surround the flask with ice and water; now add half an ounce (troy) of bromine, and agitate occasionally, until the solution has almost lost its red color; to the solution, when cold, again add a like amount of bromine, and, when the reaction is finished, add another half ounce (troy) of bromine, and so continue successively until a total of two troy ounces of bromine has been used; after each addition, permit the reaction to cease, and the liquid to cool. Now warm the flask, and when the solution has acquired a greenish color, filter; wash the excess of iron and the liberated carbon [the carbon is from the iron] with a little distilled water, and filter into the former solution. Transfer this solution at once into a clean iron dish, placed upon a sand bath, and quickly evaporate, until a portion, removed upon a glass rod, will solidify upon cooling; then remove the dry salt, immediately, into well-stoppered bottles."—*K. & L.*

Ferrous bromide is occasionally used in medicine. It is made in large quantities for preparing bromide of potassium.

Ferrous Carbonate. — Formula, $FeCO_3$. Molecular weight, 115.75.

This salt mixed with isomorphous carbonates of magnesium, calcium and manganese, occurs native as the *spathose iron ore.* It may be made by pouring a boiling solution of ferrous sulphate into a boiling solution of sodium carbonate, a *white* precipitate of $FeCO_3$ resulting. This rapidly absorbs oxygen, passes through shades of green and blue, and finally upon drying by exposure to

the atmosphere it changes to brick red or a yellowish-red and is then *ferric oxy-hydrate.* This change may be partly prevented by washing the precipitate quickly with hot distilled water, squeezing and enclosing it at once in a bladder and then drying it by exposure to the atmosphere. The green powder will effervesce with acids, carbon dioxide being liberated. It must be immediately enclosed in glass stoppered vials. This change of the carbonate into ferric oxy-hydrate is prevented, according to the pharmacopœia, by means of sugar and honey. When this mixture is reduced to pill mass consistence it is recognized by the U. S. P. as—

Pilula Ferri Carbonatis, *or pill of carbonate of iron.* This is the well-known *Vallet's pill mass.* Vallet's mass should be of a greenish-gray color and of a stiff pillular consistence. It should not be black throughout, but the surface is of this color. It dissolves with effervescence in dilute acids, carbon dioxide being evolved. If this mass be quickly dried in a warm-air closet it produces the *Ferri Carbonas Sacharata B. P.* similar to the foregoing. Usually, however, it is of a yellowish color. The so-called *sub-carbonate of iron* of the U. S. P., made by drying ferrous carbonate, is a ferric oxy-hydrate, Fe_2O_3-$Fe_2(OH)_6$. It is of a yellowish-red color and is employed extensively in medicine under the aforesaid name (sub-carbonate of iron), or is known as *precipitated carbonate of iron.* Both terms, however, are misnomers. Immense quantities of this article are now produced in the arts as incidental products. Often it is dried at high temperatures, is of a brick-red color, and has the composition $Fe_2O_3.H_2O$. This dissolves with more difficulty and should not be used in medicine.†

Ferrous Chloride.—Formula, $FeCl_2$. Molecular weight, 126.64.

This is easily prepared by the following—

Experiment.—Digest in a porcelain dish one part of iron turnings with four parts of hydrochloric acid, previously

† See ferric hydroxide.

diluted with two parts of distilled water; warm the mixture, then filter it and evaporate carefully until the bulk is reduced to the measure of that of three parts of the acid employed; then add a little hydrochloric acid and place the vessel in a cool situation to crystallize. These crystals are of a bluish color, resemble ferrous sulphate in appearance, and deliquesce in moist air, becoming first green and then yellow. They have the composition $FeCl_2 4H_2O$, are not used in medicine, but the solution of ferrous chloride is made in quantities as the first step in preparing *ferric* chloride.

Ferrous Iodide.—Formula, FeI_2. Molecular weight, 308.96.

Synonyms.—Ferri iodidum. Iodide of iron.

Iodide of iron is made by mixing, in a porcelain dish, two parts of iron turnings, sixteen parts of water, and one part of iodine, and allowing the mixture to remain in a cool place until the iodine dissolves. The dish is now to be gently warmed, and when the solution is colorless it is to be quickly filtered, then evaporated rapidly in a clean iron vessel until a small portion removed upon an iron wire will solidify. Now cool the dish and contents, heat the dish quickly, and when the iodide of iron next to the dish melts, slide the entire mass from the dish. Break it into pieces and preserve it in glass stoppered vials. Iodide of iron is a steel-gray, semi-crystalline mass and a fresh fracture presents a metallic lustre. It melts at 177° C. (350° F.) and oxidizes upon exposure to the air, iodine being liberated. Iodide of iron is occasionally employed in medicine for making pills, but the favorite mode of administration is in connection with solution of sugar which acts as a preservative, and is the pharmaceutical known as—

Syrupus Ferri Iodidi, (Syrup of iodide of iron).

This is made (U. S. P.) by forming a solution of iodide of iron from two troy ounces of iodine according to the preceding process, excepting that a smaller proportion of water is used. The solution is filtered into a sufficient

amount of simple syrup to produce twenty fluid ounces. Syrup of iodide of iron is of a pale green color. It changes to yellow, red, or brown by age. Exposure to sunlight will decolorize it if not too red. The change may be to some extent prevented by adding to it a small amount of hypophosphorous acid.

Ferrous Lactate.—Formula, $Fe2C_3H_5O_3.3H_2O$. Molecular weight, 287.36.

Synonyms.—Ferri lactas. Lactate of iron.

Lactate of iron is made by direct combination between dilute lactic acid and iron. For this purpose the acid is diluted with sixteen times its weight of water, and then digested with iron turnings or iron wire in a warm situation until the reaction ceases, when the solution is filtered. The filtrate is evaporated until a pellicle forms, and then is placed in a cool situation to crystallize. The reaction is explained as follows:—

$$\underset{\text{Lactic acid.}}{4HC_3H_5O_3.} + \underset{\text{Iron.}}{Fe_2} = \underset{\text{Ferrous lactate.}}{2Fe2C_3H_5O_3} + \underset{\text{Hydrogen.}}{2H_2}.$$

Lactate of iron is a granular crystalline mass or crust. The taste is sweetish and ferruginous. It dissolves in twelve parts of water at 212° F. If moist it decomposes upon exposure to the air, absorbs oxygen and turns red, passing into *ferric* lactate.

Ferrous Oxalate.—Formula, $FeC_2O_4.H_2O$. Molecular weight, 161 64.

Synonyms.—Ferri oxalas. Oxalate of iron.

This is made by simply displacing the sulphuric acid radical of ferrous sulphate with the radical of oxalic acid. Essentially the process of the U. S. P. is as follows:—

Dissolve two troy ounces of ferrous sulphate in thirty fluid ounces of water, and mix this with a solution of four hundred and thirty-six grains (436) of oxalic acid in fifteen fluid ounces of water. Wash the precipitate well with distilled water, and dry it in the air. This is a beautiful and simple experiment. Oxalate of iron is almost insoluble in water, therefore, this acid readily displaces the sulphuric acid from ferrous sulphate, the reaction being expressed as follows:—

$$\underset{\text{Ferrous sulphate (crystallized)}}{FeSO_4.7H_2O} + \underset{\text{Oxalic acid.}}{H_2C_2O_4.2H_2O} =$$
$$\underset{\text{Ferrous Oxalate.}}{FeC_2O_4.H_2O.} + \underset{\text{Sulphuric acid.}}{H_2SO_4} + \underset{\text{Water.}}{8H_2O.}$$

When the solutions are mixed a dense, lemon-yellow precipitate results. This must be washed until the free sulphuric acid is removed. When dry, oxalate of iron is a light, lemon-yellow crystalline powder, feeling like lycopodium when rubbed between the fingers. It is tasteless, and when calcined yields ferric oxide. With boiling solution of caustic potash, monoxide of iron (ferrous oxide) results.

Ferrous Phosphate—Formula, $Fe_3(PO_4)_2$. Molecular weight, 357.3.

Synonyms.—Ferri Phosphas. Phosphate of iron. Blue phosphate of iron.

Phosphate of iron is made according to the U. S. P. by mixing a solution of six troy ounces of phosphate of sodium in four pints of water, with a solution of five troy ounces of sulphate of iron in four pints of water; the precipitate is to be well washed with hot water and afterward dried. When the solutions are mixed, phosphate of iron forms as a white precipitate, of the composition $Fe_3(PO_4)_2$. The exposure during the drying process is accompanied by absorption of oxygen and the precipitate changes to blue. The composition may then be expressed (Stille & Maisch), $Fe_3 2PO_4.FePO_4.12H_2O$. Vivianite is a mineral which contains phosphate of iron, $Fe_3(PO_4)_2 + 8H_2O$. Phosphate of iron has a greenish blue color, is tasteless and without odor. It should not dissolve in water, nor should water yield an insoluble residue after digestion with it and subsequent evaporation. Phosphate of iron dissolves readily in dilute hydrochloric acid, the solution is yellow, and with both the red and the yellow prussiate of potash will yield blue precipitates.

Ferrous Sulphate.—Formula, $FeSO_4.7H_2O$, Molecular weight, 277 44.

Synonyms.—Ferri sulphas. Sulphate of iron. Sulphate of protoxide of iron. Green vitriol. Protosulphate of iron. Copperas.

Experiment.—Digest iron turnings in dilute sulphuric acid, until the acid is saturated. Filter this solution, add a little sulphuric acid to the filtrate and then evaporate it to crystallization, and place it in a cool situation that crystals may form.

When iron is digested with dilute sulphuric acid, the hydrogen of the acid is displaced and ferrous sulphate produced, thus:—

$$\underset{\text{Sulphuric acid.}}{2H_2SO_4} + \underset{\text{Iron.}}{Fe_2} = \underset{\text{Ferrous sulphate.}}{2FeSO_4} + \underset{\text{Hydrogen.}}{2H_2}.$$

Some contend that the reaction is not nearly so simple. Sulphate of iron is produced impure in large amounts, and sold as green vitriol. In this case it is derived from iron pyrites, or made by dissolving scrap iron in commercial dilute sulphuric acid. In either instance it contains sulphates of zinc, copper, or other metals. Sulphate of iron forms monoclinic crystals, of a bluish-green color. When made in neutral solutions these quickly oxidize and turn yellow, but the addition of an excess of sulphuric acid to the solution before crystallization renders them more stable. They effloresce in dry air, and when heated in a vacuum produce a white powder, water escaping, the residue being monohydrated ferrous sulphate, $FeSO_4.H_2O$. If the heat be continued, anhydrous ferrous sulphate $FeSO_4$, remains. Ferrous sulphate is insoluble in anhydrous alcohol and in ether, but dissolves in twice its weight of cold water, and in less than half its weight of water at 212° F. When ferrous sulphate is heated to 300° F. it fuses in its water of crystallization, and gives off six molecules of water, a white powder remaining. This is the—

Ferri Sulphas Exsiccata, or *dried sulphate of iron* of the pharmacopœia, usually contaminated with a little ferric oxysulphate, but this is of minor importance.

Ferrous sulphate forms with sulphates of the alkali metals double salts or alums. One of these—

Ammonium Ferrous Sulphate, $FeSO_4(NH_4)_2SO_4.6H_2O$, is used in photography and in certain analytical operations. Sometimes it is improperly called *ammonio-*

ferric alum. To prepare it dissolve a mixture of two parts of ferrous sulphate and one part of ammonium sulphate in its weight of hot water, and permit the solution to crystallize, evaporating it if necessary. Ammonium ferrous sulphate is in the form of clear green or bluish-green crystals. They change slightly in the air and have a specific gravity of 1.813. This salt is not employed in medicine and should not be confused with ammonio-ferric alum.

Ferrous Sulphide.—Formula, FeS. Molecular weight, 87.88.

Synonyms.—Sulphide of iron. Sulphuret of iron. Ferri sulphuretum.

This compound is made by rubbing a roll of brimstone against a bar of iron at a white heat, or by heating in a crucible alternate layers of iron turnings and flowers of sulphur. It usually contains an excess of iron, but as the compound is simply employed for making hydrogen sulphide the free iron is of no disadvantage.

Pure sulphide of iron is a yellowish, crystalline mass, but that of commerce has usually a steel color. When it is placed in dilute sulphuric or hydrochloric acid, ferrous sulphate or ferrous chloride is respectively formed and hydrogen sulphide is liberated. Thus with sulphuric acid:—

$$\underset{\text{Ferrous sulphide.}}{FeS} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Ferrous sulphate.}}{FeSO_4} + \underset{\text{Hydrogen sulphide.}}{H_2S}.$$

Ferrous sulphide should be kept in a close vessel as it oxidizes upon exposure to the air.

FERRIC SALTS.

These salts differ from the ferrous salts, inasmuch as they contain a less proportion of iron. Ferric salts precipitate blue with *ferrocyanide of potassium*, but not with ferridcyanide of potassium, although in the latter case the solution changes in color to olive green.

Ferric Acetate.—Formula, $Fe_2 6C_2H_3O_2$. Molecular weight, 464.96.

Synonym.—Acetate of iron.

This preparation can not be prepared dry, and is liable to decomposition in the form of solution. It is a favorite ferruginous preparation with many physicians, and a solution of it is officinal in the British Pharmacopœia as *tinctura ferri acetatis*. This is made by mixing two and one-half fluid ounces of solution of persulphate of iron (B. P.) with eight fluid ounces of alcohol, and adding this to a solution of two troy ounces of acetate of potassium in ten fluid ounces of alcohol, and after an hour filtering. Double decomposition ensues, sulphate of potassium is formed and solution of acetate of iron. The sulphate of potassium separates, as it is insoluble in alcohol.

According to the original process tincture of acetate of iron was made from *ferrous sulphate* and crystallized acetate of lead, double decomposition resulted, the product being sulphate of lead and *ferrous* acetate. Tincture of either ferric or ferrous acetate is of a red color, and has an inky taste.

Ferric Chloride.—Formula, Fe_2Cl_6, Molecular weight, 324.02.

Synonyms.—Ferri chloridum. Sesquichloride of iron. Chloride of iron. Perchloride of iron.

Ferric chloride is made according to the pharmacopœia as follows:—dissolve two troy ounces of iron wire in eight troy ounces of hydrochloric acid, then filter. To the filtrate add four troy ounces of hydrochloric acid, and having heated it to near the boiling point in a porcelain capsule, gradually add nitric acid until red fumes cease to follow, and the solution diluted with water will not give a blue precipitate with ferridcyanide of potassium. The solution is now to be evaporated upon a sand-bath until it weighs eight troy ounces and three hundred and sixty grains, when it is to be placed aside until it forms a crystalline mass.

When ferrous chloride in solution is mixed with nitric and hydrochloric acids and the mixture is heated, chlo-

rine is evolved from the mixed acids and combines with the ferrous chloride. Nitric oxide is also liberated and escapes, thus:—

$$6FeCl_2 + 6HCl + 2HNO_3 = 3Fe_2Cl_6 + 2NO + 4H_2O.$$

Ferrous chloride. Hydrochloric acid. Nitric acid. Ferric chloride. Nitric oxide. Water.

Ferric chloride may be made by heating iron wire in chlorine gas, or by passing chlorine gas in excess, into a solution of ferrous chloride, and subsequently evaporating the solution and crystallizing. Ferric chloride is an orange-yellow, or red mass, presenting a crystalline fracture.. It dissolves freely in alcohol, water, and ether. It absorbs moisture from the air and deliquesces. Ferric chloride in solution should not yield a blue precipitate with ferridcyanide of potassium. The officinal—

Liquor Ferri Chloridi, or *solution of chloride of iron,* is a solution of ferric chloride of specific gravity 1.355. Three troy ounces of iron are employed in preparing sixteen fluid ounces of this solution and the precipitate obtained from two fluid drachms of the solution, with excess of ammonia, when washed, dried, and ignited, will weigh 28.25 grains. This preparation is often sold on the market under the name, "*concentrated tincture of iron,*" a term derived from the fact that it is mixed with three parts of alcohol to make the officinal *tincture of chloride of iron,* or muriated tincture of iron.

Ferric Citrate. — Formula, $Fe_2 2C_6H_5O_7$. Molecular weight, 488.88.

Synonyms.—[Insoluble citrate of iron.] Ferri citras, Citrate of iron.

This is made according to the Pharmacopœia by evaporation of the officinal solution of citrate of iron, therefore, in making citrate of iron we must refer to solution of citrate of iron. The scaled compound sold on the market as citrate of iron, is usually of a composition more or less approaching the formula $Fe_2 2C_6H_5O_7$, but the uncertain formula of our Pharmacopœia is such as to permit an excess of either ferric oxide or citric acid. Solution of

ferric citrate is acid in reaction, and possesses the power to freely dissolve ferric oxide. The following process is after the plan of that of the Pharmacopœia, but more definite as to the point of saturation of the acid.

Take of solution of persulphate of iron (tersulphate of the U. S. P.) one pint, mix it with four pints of cold distilled water, and stir into the mixture dilute ammonia water until it is in slight excess. Then transfer the precipitate to a muslin strainer, and having permitted it to drain, wash it well with cold distilled water, and until the washings are nearly tasteless. Then transfer three-fourths of the precipitate to a porcelain evaporating dish and mix with it six troy ounces of citric acid. Place the dish on a steam or water-bath, and having applied heat, stir the mixture until the ferric oxide and the citric acid are both dissolved. Then place a thin layer of the solution on a plate of glass, and dry it by a gentle heat; if it separates in scales, evaporate the solution to a syrupy consistence and spread it upon glass, by means of a brush, and dry it. If it adheres to the glass, add ferric oxide gradually, testing it from time to time, that the scaling point may be determined. In case an excess of ferric oxide is present, the dried salt will separate from the glass as dust. In this case citric acid must be added to the solution.

Citrate of iron is found in market as yellowish-red scales. It is perfectly, but very slowly dissolved by cold water, readily by hot water, and very freely by dilute ammonia water. It is used in medicine for making pills, but for most purposes has been supplanted by citrate of iron and ammonium.

Ferric Citrate and Ammonium.

Synonyms.—Ferri et ammonii citras. Citrate of iron and ammonium. Soluble citrate of iron.

This is made by the pharmacopœial process, by mixing ammonia water with solution of ferric citrate, evaporating the solution, and scaling the salt from the glass. We find that, as made in this manner, the proportions vary to such an extent, owing to the uncertainty of the composition of officinal citrate of iron, that the product is often

troublesome to scale, and of varying solubilities. The following process is recommended as an improvement. Dissolve the ferric oxide from one pint of solution of persulphate of iron (tersulphate U. S. P.) by digesting it in a porcelain dish, at the temperature of 180° F., with four avoirdupois ounces of citric acid, and five fluid ounces of *solution of citrate of ammonium.* (Solution of citrate of ammonium for this purpose is made by dissolving eight avoirdupois ounces of citric acid in an excess of ammonia water, and evaporating the product until it is reduced to the measure of sixteen fluid ounces.) When solution is complete, evaporate the product to the consistence of syrup, and dry by spreading it on plates of glass and exposing it to the temperature of from 100° to 120° F.

Citrate of iron and ammonium is in garnet or dark red scales, soluble in water, and gives off ammonia when calcined. It should not deliquesce in the air, and if this is the case an excess of ammonia has been used in its preparation. In case the salt is finely divided and dusty, the oxide of iron is in excess. If it has a sour taste, and an acid reaction, citric acid is in excess. Citrate of iron and ammonium is usually prescribed under the name *citrate of iron,* or, *soluble* citrate of iron. It should be sweetish to the taste, and its solution is neutral to litmus paper. This compound, in common with the other scale preparations of iron, is of uncertain composition. It is generally considered as a mixture, or a combination, of ferric oxy-citrate with ammonio-ferric citrate.

Ferric Citrate and Quinine.

Synonyms.—Ferri et quiniæ citras. Citrate of iron and quinine.

This preparation is made according to the pharmacopœia by dissolving quinine in solution of citrate of iron, and evaporating the solution to the consistence of syrup, after which it is spread upon glass and scaled. It dissolves very slowly in water, and therefore, is replaced in commerce by a soluble salt containing ammonia. The proportion of the U. S. P. is such that fifty grains of the salt contain eight grains of quinine. The color of citrate

of iron and quinine varies according to mode of preparation, ranging from yellowish-green to reddish-brown.

Ferric Citrate and Strychnine.

Synonyms.—Ferri et strychniæ citras. Citrate of iron and strychnine.

This as made by the pharmacopœial process is simply a mixture of strychnine citrate with ammonio citrate of iron, in the proportion of about the one-hundredth part of strychnine (one per cent.) It has a bitter taste, but otherwise resembles ordinary soluble citrate of iron.

Ferric Ferrocyanide.—Formula, Fe_4FCy_3, or Fe_7Cy_{18}, or $Fe_7(CN)_{18}$.

Synonyms.— Ferri ferrocyanidum. Ferrocyanide of iron. Prussian blue. Insoluble Prussian blue. Williamson's blue.

This salt is made by pouring a solution of ferrocyanide of potassium into a solution of ferric sulphate (tersulphate U. S. P.) or ferric chloride, the ferric solution being in excess. The reaction is often expressed by the following simple equation :—

$$\underset{\text{Potassium ferrocyanide.}}{3K_4Fcy} + \underset{\text{Ferric sulphate.}}{2Fe_2(SO_4)_3} = \underset{\text{Potassium sulphate.}}{6K_2SO_4} + \underset{\text{Ferric ferrocyanide.}}{Fe_4Fcy_3},$$

Ferrocyanide of iron is a beautiful blue bulky precipitate, insoluble in water. It is tasteless, and when dried forms masses that are permeated with crevices which cause it to readily break into small pieces upon handling. The appearance of the broken lump is that of a beautiful bronze, and which yields by trituration a deep blue powder. It is insoluble in dilute mineral acids, but with solution of oxalic acid dissolves readily, and in this case is known as *soluble Prussian blue*, and is sold for making ink, or for washing purposes. This is not *true* "soluble Prussian blue," which was discovered by Berzelius and is

Ferric Potassium Ferrocyanide.—This is made by pouring a solution of ferric sulphate or ferric chloride into an excess of a solution of potassium ferrocyanide, and afterward washing the precipitate until the chloride or

sulphate of potassium is removed, and the washings turn blue. This compound has the composition

$$K_2\overset{vi}{Fe}_2(C_3N_3)_4\overset{iv}{Fe}_2.$$

It will dissolve in pure water, and the solution is precipitated by various salts. It is not used in medicine.†

Ferric Hypophosphite.—Formula, $Fe_2 6PH_2O_2$. Molecular weight, 501.08.

Synonyms.—Ferri hypophosphis. Hypophosphite of iron.

This preparation may be made by double decomposition between solution of ferric chloride or ferric sulphate, and solution of hypophosphite of sodium. If the ferric salt is free from excess of acid, and the sodium hypophosphite is previously rendered slightly acid with hypophosphorous acid, the hypophosphite of iron will precipitate when the solutions are mixed, after which it may be washed with distilled water and then dried. This is an unsatisfactory preparation, and it is customary in practice to use *ferrous hypophosphite* as a component of the popular pharmaceutical preparation, *syrup of the hypophosphites*. In this case the solution of *ferrous hypophosphite* is made by double decomposition between solution of hypophosphite of calcium and solution of ferrous sulphate. It is not dried, but mixed with the syrup in form of solution.

Ferric Nitrate.—Formula, $Fe_2 6NO_3$. Molecular weight, 483.14.

Synonyms.—Nitrate of iron. Ferri nitras. Pernitrate of iron.

† *Turnbull's Blue*, or Ferrous Ferricyanide, is the blue precipitate formed when ferridcyanide of potassium (red prussiate of potash) is added to the solution of a ferrous salt. Upon this fact we depend for the test adopted for distinguishing the ferrous salts from *ferric* salts. Thus in making perchloride of iron, (ferric chloride,) we continue the application of heat to the solution of ferrous chloride, and add nitric acid as long as the solution is precipitated blue by red prussiate of potassium. It is well to bear in mind that *ferridcyanide* (red prussiate) of potassium gives a blue precipitate with ferrous salts, but *not* with ferric salts; and that ferrocyanide (yellow prussiate) of potassium gives a blue precipitate with ferric salts, but *not* an immediate blue precipitate with ferrous salts.

This preparation is made by dissolving iron turnings in nitric acid, filtering the solution, and then adding to the filtrate an additional amount of concentrated nitric acid. Crystals will be deposited after a little time of either the composition

$$Fe_2 6NO_3 . 12H_2O, \text{ or, } Fe_2 6NO_3 . 18H_2O,$$

the water of hydration depending upon the proportion of free acid. Ferric nitrate is only used in medicine in the form of a solution, prepared according to the U.S P. in substance, as follows:—

Mix twelve fluid ounces of water with two and one-half troy ounces of iron wire, and gradually add a mixture of three troy ounces of nitric acid and six fluid ounces of water, keeping the vessel cool by means of ice water, while the reaction progresses. When effervescence ceases, filter the greenish solution, heat it in a porcelain evaporating vessel to the temperature of 130° F., and then add two additional troy ounces of nitric acid and heat it gently until effervescence ceases. Lastly, mix the product with distilled water in sufficient amount to produce thirty-six fluid ounces. When the iron is dissolved with cold dilute nitric acid, ferrous nitrate is produced ($Fe2NO_3$). When this is heated with the remainder of the nitric acid, ferric nitrate is formed at the expense of the nitric acid, nitric oxide and water being evolved. Solution of nitrate of iron is of a reddish-brown color when fresh. It is an unsatisfactory preparation, decomposing by age with deposition of insoluble oxy-nitrates.

Ferric Oxide. — Formula, Fe_2O_3. Molecular weight, 159.68.

Synonyms.—Sesquioxide of iron. Peroxide of iron.

This occurs naturally as the red hæmatite or specular iron ore, a steel-gray mineral. It is also found in other minerals in a greater or less degree of purity. It may be made artificially by igniting the ferric salt of a volatile acid. It may also be made by heating hydroxide of iron to redness. It is not employed in medicine.

Ferric Hydroxide.—Formula, $Fe_2(OH)_6$. Molecular weight, 213.56.

Synonyms.—Ferri oxidum hydratum. Hydrated oxide of iron. Hydrate of iron. Hydrated sesquioxide of iron. Hydrated peroxide of iron. (Antidote to arsenic.)

Mix either solution of ferric chloride or solution of ferric sulphate, with four times its bulk of cold distilled water, and add to this, with constant stirring, ammonia water until the latter is in slight excess. Drain the precipitate on a muslin strainer and wash it well with cold distilled water. This forms the well known—

Antidote to Arsenic.†—Freshly prepared ferric hydroxide is very soluble in acids. When it is kept for any length of time, half its water of hydration separates, and a crystalline formation results; this is indifferent to the action of acids, and is not the antidote to arsenic. When ferric hydroxide is mixed with solution of arsenious acid, insoluble *arsenite* of iron results. Undoubtedly only the outside of the particles of undissolved arsenious acid (arsenic) is coated with this insoluble compound, therefore, it must be remembered that it is necessary to evacuate the stomach as soon as possible by means of an emetic or a stomach pump. Freshly precipitated ferric hydroxide is a gelatinous substance and has a dark brownish color, appearing red and transparent in thin layer, as between two plates of glass. It is tasteless and odorless and by

† The U. S. P. directs that solution of tersulphate of iron be used in making this preparation. Solution of tersulphate of iron, U. S P. is in reality solution of *ferric* sulphate, but this is very seldom employed in medicine, being used exclusively in making other salts of iron. It is well to remember that in case of an emergency, either solution of tersulphate of iron or subsulphate of iron, (Monsel's solution), or perchloride of iron (ferric chloride), or the tincture of muriate of iron, will answer for this purpose. If *ammonia* water can not be obtained, use either bicarbonate of sodium or solution of potash, or caustic soda (concentrated lye) to precipitate the hydroxide. Do not wait for the precipitate to drain, *squeeze it in a muslin strainer, mix the magma with water and squeeze again, then administer it.* Or, mix solution of tersulphate of iron (persulphate) with twice its bulk of water, and triturate the mixture with excess of a paste made of calcined magnesia and water. This may be administered at once and in substance.

age turns brick-red, assumes the aforementioned crystalline form, and shrinks to one-third the original bulk, water overlying. Ferric hydroxide dissolves freely in solution of ferric chloride, forming a dark red liquid of a styptic taste. If this liquid be submitted to dialysis, the crystalloids pass through the membrane and basic oxychlorides of iron remain within it. At last, only a small proportion of hydrochloric acid (or chlorides) is present within the dialyzer, and the solution is almost tasteless. This is the *so-called* solution of oxide of iron, in reality a solution of oxychloride of iron.† Hydroxide of iron may be dried, and when the residue is powdered it presents a dark brown appearance, and has the composition—

$$Fe_2O_3.Fe_2(OH)_6.$$

This substance is also formed when hydroxide of iron is allowed to freeze, and it is found native.

Ferroso-Ferric Oxide. — Formula, Fe_3O_4, or $FeO.Fe_2O_3$. Molecular weight, 231 54.

Synonyms. — Ferri oxidum magneticum. Magnetic oxide of iron. Black oxide of iron. Black hydrate of iron. Ferroso-ferric hydrate.

This article is found native, and is known as *magnetic iron ore.* It is made artificially, the process of the British Pharmacopœia being essentially as follows:—

Dissolve two troy ounces of ferrous sulphate in two pints of water, and mix the solution with five and one-half fluid ounces of solution of persulphate of iron, B. P. (tersulphate U. S. P. will answer). Mix this with four pints of solution of caustic soda, specific gravity 1.071, and bring the mixture to a boil. Then permit it to cool, stirring occasionally for two hours, drain the precipitate on a muslin strainer, and wash it until the washings pass

† This preparation is known on the market as *dialyzed iron* and is often made by adding ammonia water to solution of ferric chloride until a little of the ferric hydroxide produced refuses to dissolve. This solution is then dialyzed, chloride of ammonium passing through the membrane and oxy-chloride of iron remaining within. If the process be conducted beyond a certain point the contents of the dialyzer will gelatinize.

tasteless. Then dry the magma. Ferroso-ferric oxide is a black-brown powder, tasteless, and soluble in hydrochloric acid without effervescence. It is capable of being attracted by the magnet. This preparation, more or less impure, has been long employed in medicine, under the name *Æthiops Martialis.* The chemistry of ferroso-ferric oxide may be clearly brought to mind by considering the fact that metallic iron forms two basic oxides. Iron monoxide† or *ferrous* oxide, FeO, and iron sesquioxide or ferric oxide, Fe_2O_3, which unite to form *ferroso-ferric oxide,* thus:—$FeO.+Fe_2O_3=Fe_3O_4$.

Ferric Pyrophosphate.—Formula, $Fe_4 3P_2O_7.9H_2O$. Molecular weight, 906.16.

Synonyms.—Ferri pyrophosphas. Pyrophosphate of iron.

This substance results as a white precipitate, when cold solution of pyrophosphate of sodium is mixed with solution of persulphate of iron (tersulphate U. S. P.), thus:

$$\underset{\text{Sodium pyrophosphate.}}{3Na_4P_2O_7} + \underset{\text{Ferric sulphate.}}{2Fe_2 3SO_4} = \underset{\text{Ferric pyrophosphate.}}{Fe_4 3P_2O_7} + \underset{\text{Sodium sulphate.}}{6Na_2SO_4}.$$

Pyrophosphate of iron is insoluble in water, but readily dissolves in solution of citrate of ammonium, forming a greenish-yellow solution. This evaporated to a syrupy consistence and spread upon glass, separates when dried in the form of apple-green scales. These scales are recognized by the Pharmacopœia as *Ferri Pyrophosphas.* Scaled, so-called pyrophosphate of iron, is thus a combination of ammonia, citric acid, and pyrophosphate of iron. The actual chemical constitution of the preparation is undecided. Pyrophosphate of iron is quite soluble when fresh; but becomes less soluble by age. Sunlight decomposes it. This is the form of iron employed so extensively in the class of pharmaceutical preparations known as *elixirs.*

† Ferrous oxide is not used in medicine. It may be made by adding oxalate of iron to a boiling solution of caustic potash. It is a black powder and absorbs oxygen from the air.

Ferric Sulphate. — Formula, $Fe_2(SO_4)_3$. Molecular weight, 399.26.

Synonyms.—Persulphate of iron. Tersulphate of iron.

This salt is formed when ferrous sulphate is dissolved in water, with excess of sulphuric acid, nitric acid being then added to the heated mixture as long as red fumes escape. The reaction is expressed as follows :—

$$6FeSO_4 + 3H_2SO_4 + 2HNO_3 = 3Fe_2(SO_4)_3 + 4H_2O + 2NO.$$

Ferrous sulphate. Sulphuric acid. Nitric acid. Ferric sulphate. Water. Nitric oxide.

Persulphate of iron is the form of ferric salt directed by the U. S. P. as the basis for ferric hydroxide and the scaled salts of iron. It is known in solution as—

Liquor Ferri Tersulphatis, or *solution of tersulphate of iron,* and as such is recognized by our Pharmacopœia. The mode of preparing the solution is essentially as follows :—

Mix one troy ounce and three hundred and sixty grains of nitric acid with two troy ounces and sixty grains of sulphuric acid, and half a pint of water, in a capacious porcelain capsule, and heat the mixture to the boiling point. Add gradually twelve troy ounces of ferrous sulphate, heat until excess of nitric acid is expelled, cool and then add water until the contents of the dish will measure twenty-four fluid ounces. Solution of *tersulphate* of iron, U. S. P., is identical in chemical composition with solution of *persulphate* of iron, B. P. The strength of the British solution (specific gravity 1.441) is a little greater than that of our own (specific gravity 1.320). The color of this preparation is brown. It has no odor and is very acid and astringent to the taste. If it be diluted with water, it changes to a reddish color. The foregoing preparation must not be confused with the well known styptic,

Monsel's Solution (Liquor Ferri Sub-sulphatis, U. S. P. often called solution of persulphate of iron.)

This preparation is made in the same manner as solution of tersulphate of iron, but the sulphuric acid is not in sufficient amount for the production of a true ferric

sulphate, therefore, an oxy-sulphate results. Monsel's solution, or solution of sub-sulphate of iron† should always be designated by one of the terms just named, and *not* as persulphate of iron. It is of a dark brown-red color, and of specific gravity 1.552. It is very astringent, and is not acid to the taste like tersulphate of iron. The use is as a styptic, and as such it is among the most valuable. There are many basic ferric sulphates known, but the above only is of use in medicine. If Monsel's solution be placed in an open dish, and exposed to the atmosphere, especially near escaping steam, a light-yellow mass will result, which when powdered is known as *Monsel's salt.* It is employed as a styptic, and is a basic ferric oxy-sulphate.

Ferric Sulphate and Ammonium—Formula $(NH_4)_2Fe_2(SO_4)_4.24H_2O$. Molecular weight, 962.14.

Synonyms. — Ferri et ammonii sulphas. Sulphate of iron and ammonium. Ammonio-ferric alum.

Dissolve four and one-half ($4\frac{1}{2}$) troy ounces of sulphate of ammonium in two pints of solution of persulphate of iron (tersulphate U. S. P.) by the aid of heat. Place the solution in a cool place to crystallize. This is a double salt of ammonium and ferric sulphates, and as such may be represented by $(NH_4)_2SO_4.Fe_2(SO_4)_3.24H_2O$. It crystallizes in regular octohedrons, is astringent, and from analogy is called *alum.* According to the writer's experience it is unnecessary to heat the solution of the two sulphates above 150° F. A temperature of 212° F., decomposes the salt into the two sulphates, but they reunite as the solution cools. Sulphate of iron and ammonium is of a pink color when warm, of a beautiful bluish purple when cold and first crystallized. It will keep if closely corked, or if preserved in an atmosphere which is saturated with moisture. If a large crystal be suspended in

† The meaning of the prefix *sub* (under) here is only with reference to the highest sulphate, the ferric sulphate. Sub-sulphate of iron is intermediate between ferrous sulphate, $FeSO_4$, and ferric sulphate $Fe_2(SO_4)_3$.

a close glass over water, it will part with its water of crystallization when the weather is warm, and this saturated with the salt, will fall in drops from the lower surface. The crystal absorbs moisture again when cool and thus solution will alternately follow crystallization, and the crystal will gradually be transferred to the vessel below. Sulphate of iron and ammonium should not be exposed to the atmosphere, as it gradually decomposes under such conditions, absorbs oxygen, parts with its water of crystallization, and a mixture of ammonium sulphate and ferric oxy-sulphates remains.

Ferric and Ammonium Tartrate.

Synonyms.—Ferri et ammonii tartras. Tartrate of iron and ammonium.

This is one of the scaled salts of iron and is officinal in the U. S. P., but may well be discontinued as it is seldom prescribed. It is prepared by neutralizing a given amount of tartaric acid in solution with ammonium carbonate (ammonia water is better) and afterward dissolving in this solution as much tartaric acid as was originally used. This is then saturated with ferric hydroxide, and the solution is evaporated to a syrupy consistence and scaled from plates of glass. It is in garnet red scales, of a sweetish taste, and when heated decomposes, evolving ammonia.

Ferric and Potassium Tartrate.

Synonyms.—Ferri et potassii tartras. Tartrate of iron and potassium. Ferri tartarisatum. Ferro-tartrate of potassium. Ferri potassio-tartras.

This is one of the scale salts of iron, of somewhat doubtful composition. It is made by digesting ferric hydroxide with potassium bitartrate (cream of tartar), filtering the solution, evaporating the filtrate to the consistence of a syrup and then scaling it from glass in the manner explained under citrate of iron. Tartrate of iron and potassium is usually in brown scales, of a sweetish taste. It slowly dissolves in cold water An improved salt is made by the addition of a little ammonia water,

which renders the salt soluble in less than its weight of cold water.

Tests for Iron.—Ferrous salts in slightly acid solution produce white precipitates with ferrocyanide of potassium. The precipitate rapidly absorbs oxygen from the air and turns blue. Ferrous salts are prone to absorb oxygen from the air and pass into *ferric salts*, therefore, the precipitate is usually of a bluish white from the presence of ferric salts. If the solution of a ferrous salt be mixed with a solution of ferricyanide of potassium (red prussiate of potash) a deep blue precipitate at once follows, insoluble in hydrochloric acid.

Ferric salts form blue precipitates with ferrocyanide of potassium (yellow prussiate of potash). If the solution of ferric salt is alkaline it must be made acid with acetic acid. If it contain free mineral acid, add an excess of an alkaline acetate before applying the test. Ferric salts in neutral or moderately acid solution strike a blood-red color with potassium sulphocyanide. This test is extremely delicate. An alkaline acetate in excess prevents the reaction.

MANGANESE.

Symbol, Mn. Atomic weight, 54.8.

Scheele, in 1774, investigated the compounds of manganese. Gahn afterward isolated the metal. Manganese occurs native as the dioxide MnO_2, or black oxide of manganese, as it is usually termed. It is present in many other minerals. Manganese may be made by several similar processes, including electrolysis of the chloride. Manganese is gray and appears like cast-iron. It is hard, brittle and of the specific gravity 8.0. It unites rapidly with oxygen, decomposes water, liberating hydrogen, and like potassium and sodium must be kept under naptha or benzine. It melts at a white heat, and when mixed with iron and melted and cast, it does not oxidize, or at least very slowly. A few salts of manganese are used in medicine to a slight extent. Chemists recognize two classes of salts of manganese, viz., *manganous compounds* and *manganic compounds.*

Manganous Chloride.—Formula, $MnCl_2$. Molecular weight, 125.54.

Synonyms.—Manganesii chloridum. Chloride of manganese. Dichloride of manganese.

This salt is obtained in large amounts where chlorine gas is made by means of black oxide of manganese. In this case the impure chloride of manganese which results from the action of hydrochloric acid upon the black oxide, is purified as follows:—

Evaporate the solution, until it is free from excess of hydrochloric acid. Remove one-fourth of it and add to it an excess of solution of carbonate of sodium. Wash the precipitate well with water, and mix it with the remaining three fourths of original solution. Boil the mixture until solution of ferrocyanide of potassium fails to strike a blue color with a filtered portion of it, (the iron is now precipitated). Filter, and evaporate the filtrate to crystallization, and crystallize in a cold situation. The crystals should be purified by recrystallization from distilled water. Chloride of manganese crystallizes in rose-colored crystals of the monoclinic system, which have four molecules of water of crystallization, $MnCl_2.4H_2O$. They fuse in their water of crystallization at less than the boiling point of water, and lose two molecules of water when placed over sulphuric acid. The salt is soluble in one-sixth its weight of water at 60° C. (140° F.) and unites with chlorides of the alkali metals to form double salts. Chloride of manganese is the preparation from which the other salts of this element are generally derived, carbonate being first prepared as follows:†—

Manganous Carbonate.—Formula, $MnCO_3$. Molecular weight, 114.65.

This is obtained as a white precipitate by mixing solution of carbonate of sodium with solution of either chlo-

† Manganic chloride may be obtained in solution by dissolving manganic hydroxide, $Mn_2O_2(OH)_2$, in cold hydrochloric acid. It has the composition Mn_2Cl_6, and when the solution is heated decomposes, chlorine being evolved and manganous chloride remaining, thus:—$Mn_2Cl_6 = 2MnCl_2 + Cl_2$.

ride of manganese or solution of sulphate of manganese. It must be washed with hot distilled water and quickly dried. When dry it has a buff color, is tasteless, and dissolves in 8000 parts of water (R. & S.). It readily unites with acids to form salts, and is dissolved freely by carbonic acid water.

Manganous Acetate—Formula, $Mn2C_2H_3O_2+2H_2O$.

This is made by adding manganous carbonate to acetic acid as long as effervescence continues, filtering the solution, evaporating it to the consistence of a syrupy liquid, adding a little acetic acid and crystallizing in a cold situation. The reaction is expressed as follows:—

$$\underset{\text{Manganous carbonate.}}{MnCO_3} + \underset{\text{Acetic acid.}}{2HC_2H_3O_2} = \underset{\text{Manganous acetate.}}{Mn2C_2H_3O_2} + \underset{\text{Carbonic acid.}}{H_2CO_3}.$$

Acetate of manganese is in the form of deep rose-colored crystals, very soluble in water.

Manganous Iodide, MnI_2+4H_2O, or *iodide of manganese*, is made by saturating solution of hydriodic acid with manganous carbonate, thus:—

$$\underset{\text{Manganous carbonate.}}{MnCO_3} + \underset{\text{Hydriodic acid.}}{2HI} = \underset{\text{Manganous iodide.}}{MnI_2} + \underset{\text{Carbonic acid}}{H_2CO_3}.$$

The salt may be obtained in the form of crystals, but it is very deliquescent and is unstable. A *syrup* of iodide of manganese is used in medicine.

Manganous Phosphate, or *monohydrogen manganous phosphate*, $HMnPO_4.3H_2O$., is produced when a solution of ordinary phosphate of sodium is precipitated by means of a solution of sulphate, or of chloride of manganese, thus:—

$$\underset{\text{Sodium orthophosphate.}}{HNa_2PO_4} + \underset{\text{Manganese sulphate.}}{MnSO_4} = \underset{\text{Manganous phosphate.}}{HMnPO_4} + \underset{\text{Sodium sulphate.}}{Na_2SO_4}.$$

In addition to the above we have normal manganous orthophosphate, $Mn_3(PO_4)_2+7H_2O$, and dihydrogen manganous phosphate, $H_4Mn(PO_4)_2.2H_2O$. Phosphate of manganese is very seldom employed in medicine.

Manganous Sulphate, $MnSO_4$, or *sulphate of manganese,* is most readily prepared from the carbonate of manganese by a process similar to that we have given for making acetate of manganese, although it can be made directly from black oxide of manganese and sulphuric acid. To prepare it from carbonate of manganese, add carbonate of manganese to dilute sulphuric acid as long as effervescence continues. Filter the solution, evaporate it to the consistence of a syrup, and place the solution in a shallow vessel, in a cool place to crystallize. It will require several days before the crystallization is complete. The reaction is explained by the equation—

$$\underset{\text{Manganous carbonate.}}{MnCO_3} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Manganous sulphate.}}{MnSO_4} + \underset{\text{Carbonic acid.}}{H_2CO_3}.$$

Manganous sulphate varies in properties according to the water of crystallization present. If the salt is crystallized at a less temperature than 6° C. (42.8° F.) it will have the composition $MnSO_4.7H_2O$; between 7° and 10° C. (44.6° and 50° F.) it is represented by $MnSO_4.5H_2O$, and at from 20° to 30° C. (68° to 86° F.) by $MnSO_4.3H_2O$.

Manganous sulphate is rose colored, and dissolves freely in water (we refer to the common salt of commerce, $MnSO_4.3H_2O$). It is the salt of manganese most used in medicine (none are in general use.) It will be necessary to study the oxides of manganese and the oxy-acids, to clearly understand the nature of the *permanganates,* the manganates, the manganous, and the manganic salts.

Oxides of Manganese.

These may be named as follows:

Manganous oxide, or manganese monoxide, ...MnO.
Mangano-manganic oxide, or red manganese oxide Mn_3O_4.
Manganic oxide, or manganese sesquioxide.....Mn_2O_3.
Manganese dioxide, manganese peroxide, or black oxide of manganese..... ....MnO_2.
Manganese heptoxide..... ..... . Mn_2O_7.

The first of these oxides is a powerful base (MnO); the last (Mn_2O_7) an acid producer which will form permanganic acid ($HMnO_4$) when brought into contact with water.

The fourth (MnO_2) also acts occasionally as a weak acid, but generally the *three* intermediate oxides are feebly basic. There are several poorly defined oxides beside those we have named, but it is unnecessary for us to consider them. Indeed, many believe such to be mere combinations of the foregoing, and in like manner define red manganese oxide, Mn_3O_4. as a compound of manganese monoxide (MnO) and manganic oxide (Mn_2O_3), thus:—

$$MnO + Mn_2O_3 = Mn_3O_4.$$

Manganous oxide, (Manganese monoxide,) MnO.

This is produced when a higher oxide of manganese is heated to redness in an atmosphere of hydrogen. It is a grayish-green powder of specific gravity 5.09. This oxide is of interest as it gives rise to the series of compounds known as *manganous compounds* (previously considered). It is a powerful base. When solutions of the manganous salts are precipitated with an alkali, a white hydrate falls. This is known as *manganous* hydrate $Mn(OH)_2$, and has the composition of a molecule of manganous oxide united with a molecule of water, thus:—

$$MnO + H_2O = Mn(OH)_2.$$

The reaction which results in the decomposition of manganous chloride by means of potassium hydroxide is represented as follows:—

$MnCl_2$	+	$2KOH$	=	$Mn(OH)_2$	+	$2KCl$.
Manganous chloride.		Potassium hydroxide.		Manganous hydrate.		Potassium chloride.

Manganous hydrate oxidizes rapidly upon exposure to the air and turns brown, *manganic* hydroxide, $Mn_2O_2(OH)_2$, being first formed.

Manganic Oxide, (Sesqui-oxide of manganese,) Mn_2O_3.

This oxide is typical of a series of manganese salts known as *manganic compounds*, none of use in medicine. It is found native as the mineral *braunite*. Sesquioxide of manganese is produced when the other oxides of manganese are ignited in oxygen gas. The contents of the retorts which are employed in preparing oxygen by heating manganese dioxide to redness, consists mainly of this

Manganous Acid —Manganate of potassium is believed to be a salt of *manganous acid*, H_2MnO_4, which has never been obtained in a free state. Although hypothetical, it is accepted that it gives rise to a series of well known oxy-salts, the *manganates*, thus :—

Manganous acid	H_2MnO_4.
Potassium manganate	K_2MnO_4.
Sodium manganate	Na_2MnO_4.
Barium manganate	$BaMnO_4$.

Manganese Heptoxide, or *permanganic anhydrid*, Mn_2O_7, is a very unstable compound. It is made by dissolving perfectly pure permanganate of potassium in very strong sulphuric acid, when it separates and settles as an oily liquid. It readily evolves oxygen, and decomposes with explosive violence when heated, or when it is brought into contact with many organic substances. This should not be made for experiment, on account of the danger of explosions.

Permanganic Acid, $HMnO_4$.—When a solution of manganate of potassium is boiled, decomposition ensues and a red solution is produced dependent upon the formation of potassium permanganate. At the same time caustic potash and hydrated dioxide of manganese are formed, thus :—

$$\underset{\text{Potassium manganate.}}{3K_2MnO_4} + \underset{\text{Water.}}{3H_2O} = \underset{\text{Potassium permanganate.}}{2KMnO_4} + \underset{\text{Caustic potash.}}{4KOH} + \underset{\text{Hydrated manganese dioxide.}}{MnO_2,H_2O}.$$

Potassium permanganate† may be considered a compound which is derived from displacement of the hydrogen in permanganic acid. A series of permanganates are known ; thus, as examples :—

Permanganic acid	$HMnO_4$.
Potassium permanganate	$KMnO_4$,
Sodium permanganate	$NaMnO_4$.
Barium permanganate	$Ba(MnO_4)_2$.

Permanganic acid, in aqueous solution, may be obtained from *barium permanganate* by decomposition with

† See potassium permanganate, under salts of potassium, for practical preparation.

sulphuric acid. It is a dark-red liquid, a powerful oxidizer, and decomposes even by the action of light, oxygen being evolved.

Tests for Manganese.—Mix compounds of manganese with sodium carbonate and a small amount of potassium nitrate. Fuse the mixture upon platinum foil in the outer blowpipe flame. Sodium manganate results, of a green color. Dissolve the same in a small amount of distilled water, a green solution results. Add excess of acetic acid, the solution changes to red. If there be organic matter present brown flocculi separate upon standing and the solution becomes colorless. If limestone be examined in this manner a little borax must be added previous to fusing the mixture. Salts of manganese in solution are characterized by giving flesh-colored precipitates with sulphide of ammonium. Hydroxides of potassium and of sodium give white precipitates with solutions of manganese salts which change to brown upon exposure.

NICKEL.

Symbol, Ni. Atomic weight, 58.6.

The ores of this metal have been distinguished from copper since 1694. Bergman, in 1774, identified the metal, and supported the results of Cronstedt's investigations made some thirteen years previously. Thus Cronstedt is the discoverer of the element. Nickel is extracted from the ore by calcination, and also by a wet method. It is always impure as found in commerce, being contaminated with iron, copper, carbon, and, excepting that from New Caledonia, cobalt. The pure metal may be made by reducing the chloride in an atmosphere of hydrogen gas, or by igniting oxalate of nickel out of contact with the air. Nickel is silver or grayish-white and very hard. It has the specific gravity 8.9 (R. & S.), is malleable, magnetic and permanent in the air. It unites with oxygen at redness and dissolves in dilute nitric, sulphuric and hydrochloric acids. It is useful as an alloy in making certain cheap coin, and in preparing German

silver. Nickel forms two oxides, a *monoxide*, NiO, and a *sesquioxide* or *peroxide*, Ni_2O_3. Neither is of use. There are many salts of nickel, one only of general use in the arts, and this (sulphate) has been recommended in medicine, but it is very seldom employed as a therapeutical agent.

Nickel Sulphate.—Formula, $NiSO_4$.

This salt is made by dissolving carbonate of nickel in dilute sulphuric acid, evaporating the solution and crystallizing. Sulphate of nickel is a beautiful emerald-green salt, crystallizing with either six or seven molecules of water according as the solution is acid or neutral. Sulphate of nickel is used mainly in electro-plating, for which purpose it is dissolved in water acidulated with sulphuric acid, and then made into a double salt with ammonium sulphate. This, *ammonium nickel sulphate* has the composition $(NH_4)_2SO_4.NiSO_4+6H_2O$.

CHROMIUM GROUP OF METALS.

CHROMIUM.	TUNGSTEN.
MOLYBDENUM.	URANIUM.

This group of metals forms a series of trioxides which give rise to oxy-acids, and these in turn produce an interesting line of salts. Chromium alone is of value in a medicinal sense, but molybdenum yields compounds which are very useful as tests for alkaloids.

CHROMIUM.

Symbol, Cr. Atomic weight, 52.4.

This metal was discovered in 1797 by Vanquelin. He gave it the name *chromium* with reference to the yellow color of its salts. Chromium does not occur native, but may be obtained from the oxide or chloride by reduction at a high temperature with sugar. Chromium has a grayish-white color, the specific gravity 6.8 (Wœhler), or 7.3 (Bunsen). It does not oxidize in the air, but forms chromic oxide (Cr_2O_3) at redness. Nitric acid does not

affect it; warm dilute sulphuric acid dissolves it, hydrogen escaping; dilute hydrochloric acid dissolves it quickly. The oxides and the salts of chromium are of exceeding interest to chemists, many are of great value in the arts. Very few are used in medicine.

Chromium and Oxygen.

Chromium forms two oxides of a basic nature, and one which is an acid former. Chromium monoxide or chromous oxide (hypothetical) CrO. Chromium sesquioxide or chromic oxide (green oxide of chromium) Cr_2O_3. Chromium trioxide, or chromium anhydrid (usually called chromic acid), CrO_3.

Chromium Trioxide. — Formula, CrO_3. Molecular weight, 100.28.

Synonyms.—Chromium anhydrid. Acidum chromicum. Chromic acid.

This is the only oxide of interest to the physician. It is made by decomposing solution of bichromate of potassium with sulphuric acid, the chromium anhydrid separating as brilliant scarlet or red needles (see chromic acid). Chromium trioxide is an energetic oxidizing agent, many organic and inorganic bodies reducing it quickly, chromic oxide, Cr_2O_3, being formed. When chromium trioxide is dissolved in water a yellow solution results. This is in reality a solution of *chromic acid*, H_2CrO_4, and is produced by the combination of a molecule of water with a molecule of chromium trioxide, thus:—

$$CrO_3 + H_2O = H_2CrO_4.$$

This is typical of a series of salts known as—

Chromates, of which we may mention as examples,

Normal potassium chromate..... K_2CrO_4.
Normal ammonium chromate..... $(NH_4)_2CrO_4$.
Barium chromate..... $BaCrO_4$.
Lead chromate..... $PbCrO_4$.

The lead chromate is employed in the arts as a yellow pigment (chrome yellow). The common bichromates, or *dichromates*, are a class of salts represented by $M_2Cr_2O_7$.

They have the constitution of a molecule of chromium trioxide and a molecule of a neutral chromate, thus:—

Potassium chromate .....$K_2CrO_4 + CrO_3 = K_2Cr_2O_7$.
Ammonium " $(NH_4)_2CrO_4 + CrO_3 = (NH_4)_2Cr_2O_7$.
Sodium chromate.....$Na_2CrO_4 + CrO_3 = Na_2Cr_2O_7$.

Or, we may call the neutral chromates compounds formed from hydroxides with one molecule of chromium trioxide, thus:— $2KOH + CrO_3 = K_2CrO_4 + H_2O$; and the bichromates compounds formed from hydroxides with two molecules of chromium trioxide, thus:—

$$2KOH + 2CrO_3 = K_2Cr_2O_7 + H_2O.$$

Bichromate of potassium is employed in making officinal chromic acid, and has been described under potassium salts. It is not necessary for us to consume space with the other preparations of chromium.

MOLYBDENUM.

Symbol, Mo. Atomic weight, 95.6.

Scheele pointed out the fact that sulphide of molybdenum differed from graphite. It seems, however, that Hjelm first prepared the metal. Molybdenum is made by reduction of the trioxide in hydrogen gas, or by heating a chloride of molybdenum in an atmosphere of hydrogen. Molybdenum is silver white, of specific gravity 8.6, and refuses to melt at the highest heat of a wind furnace. It is very hard, does not oxidize at ordinary temperatures, dissolves in nitric acid and in hot sulphuric acid, but not in dilute sulphuric nor hydrochloric acid. It forms a number of oxides, oxy-chlorides and chlorides; also an acid which is known as molybdic acid H_2MoO_4, and this is typical of a series of salts known as *molybdates*, thus:—

H_2MoO_4	+	K_2CO_3	=	K_2MoO_4	+	H_2CO_3.
Molybdic acid.		Potassium carbonate.		Potassium molybdate.		Carbonic acid.

None of the compounds of molybdenum are used as medicines, but molybdic acid, molybdate of ammonium $(NH_4)_2MoO_4$, and phospho-molybdic acid are of great value as a test for alkaloids.

TUNGSTEN (*Wolfram.*)

Symbol, W. Atomic weight, 184.0.

This metal is only found in a few minerals. None of its compounds are employed in medicine. It is yielded most abundantly by the ore *wolfram*. When a mixture of tungsten trioxide (WO_3) and carbon is calcined in a close crucible, the metal is liberated, but as yet it has never been obtained in mass. The powder has a metallic lustre, specific gravity 19.120, burns at a red heat in the air and oxidizes rapidly in nitro-muriatic acid. It is said to harden steel when in small proportion. The symbol W. was derived from the name *wolfram* applied to the common ore of the metal.

URANIUM.

Symbol, U. Atomic weight, 240.0.

Klaproth named this element after the planet *Uranus*, in 1789. He did not obtain the metal, however, as he supposed. Peligot isolated it from a mixture of uranous chloride and either potassium or sodium, by calcination. Uranium has a specific gravity of 18.33, is hard, and of a whitish nickel-like color. It is of no use in the arts, and none of its compounds are employed in medicine.

TIN GROUP OF METALS.

TIN. THORIUM.
TITANIUM. ZIRCONIUM.

Of this group tin alone is of value in the arts, and at one time powdered tin was employed in medicine. The metals of this group form dioxides which correspond with dioxide of silicon, (SiO_2.) They resemble this non-metallic element, furthermore, from the fact that their *fluorides* form double salts with the fluorides of other metals, and these are isomorphous with the silico-fluorides.

TIN (*Stannum.*)

Symbol, Sn. Atomic weight, 117.8.

Tin has been known since a very early day, and is found in many parts of the world, but not plentifully outside of a few localities. The *dioxide* of tin or the mineral *cassiterite* yields the tin of commerce. Common tin contains arsenic, bismuth, antimony, zinc, copper, and other metals. Pure tin is white, brilliant, harder than lead, and of the specific gravity 7.293. When a bar of tin is bent it emits a crackling noise which is known as "the cry of tin." Tin may be rolled or hammered, and is employed in making a foil which is extensively used in various branches of the arts.† Tinned vessels of copper, iron and brass are in common use. Tin forms several alloys with other metals; as such we name pewter, gun-metal, bronze and bell-metal. It also combines with mercury to form an amalgam which is valuable as a gilding for glass.

Tin Powder.

This is the only form of tin that has ever had any reputation as a remedy, and it is now obsolete. It was used once for the expulsion of tape-worms. Powdered tin is made by pouring melted tin into a mortar and trituating rapidly until it cools. Tin is often dissolved in dilute hydrochloric acid, for this purpose it is advisable to *granulate* the metal. This can be done by pouring the molten metal, in a small stream, into cold water.

Oxides.

Tin forms two oxides, monoxide or *stannous* oxide, SnO, and dioxide or *stannic* oxide, SnO_2. Stannous oxide is of little importance, but the stannic oxide in an impure form is the natural ore of tin. These two oxides are typical of two classes of salts known severally as *stannous* salts and *stannic* salts. Of these one is used extensively as a mordant in dyeing under the name *tin salt*. This—

† Ordinary tin-foil is largely composed of lead. It is made by rolling a piece of lead between two pieces of tin, thus making a foil with a tin surface.

Stannous Chloride, or *dichloride of tin,* is made by dissolving metallic tin in hydrochloric acid, and evaporating the solution to crystallization. The crystals are colorless, needle-like, and have the composition, $SnCl_2$-$2H_2O$. In commerce a solution of stannous chloride is known under the name, "*solution of tin,*" and is used for the same purpose as the crystals.

Tin forms many compounds of interest to chemists, but which we are not called upon to review. The disulphide of tin (SnS_2) is made by heating a mixture of sal-ammoniac, sulphur, and an amalgam of tin. It has a bronze color, and is known as *mosaic gold.*

The soluble salts of tin are poisonous. In cases of poisoning, vomit the patient, use the stomach pump, and administer mucilaginous and albuminous drinks. Carbonate of ammonium has also been recommended.

Tests for Tin.—If a tin compound be held in the reducing flame of the blow-pipe upon a carbon pencil, reduction takes place, and a bead of metallic tin remains. This dissolves in hydrochloric acid, and the solution precipitates red with selenious acid. With mercuric chloride a white precipitate is at first formed, and afterward a black precipitate. *Stannous salts* give brown precipitates with sulphide of hydrogen, which are insoluble in fresh, colorless sulphide of ammonium. *Stannic salts* give yellow precipitates with the same reagent, and these dissolve in ammonium sulphide. If metallic tin be acted upon by nitric acid a white powder (SnO_2) results.

TITANIUM.

Symbol, Ti. Atomic weight, 48.0.

This metal is of no importance in medicine, and is not used in the arts. It was discovered by Klaproth in 1795. It is very rare, and occurs in nature as a dioxide. It possesses a metallic lustre, and an iron-gray color. It will burn in the air, forming dioxide, contaminated with a little *nitride.* It dissolves in warm hydrochloric acid, hydrogen being evolved. Dilute sulphuric and other acids also affect it.

Titanium is remarkable for the fact that it combines at high temperatures with nitrogen. Several compounds of titanium and nitrogen are known (nitrides).

THORIUM.

Symbol. Th. Atomic weight, 231.5.

Thorium is found in several minerals which contain metals of the cerium group. It was discovered as an earth by Berzelius in 1828, the name *thoria* being given it.

Thorium is a gray powder capable of being polished. The specific gravity is 7.6, or 7.7. It dissolves in nitric acid readily, and burns in the air if heated. Thorium is obtained from the chloride by heating it in a crucible with metallic potassium or metallic sodium. It forms an oxide of the composition ThO_2 (thorium oxide or thoria), which gives rise to a line of salts of no general interest.

ZIRCONIUM.

Symbol. Zr. Atomic weight, 90.0

Zirconium was discovered as an earth and named by Klaproth in 1789. Berzelius first isolated the metal. It is very rare, and the crystalline form only unites with oxygen in the heat of an oxy-hydrogen flame. The amorphous form, upon the contrary, burns brightly when simply warmed in the air. It oxidizes in nitro-muriatic acids, dissolves in hydrofluoric acid, but common acids do not affect it, even if heated. Zirconium forms one oxide, ZrO_2, which is known as *zirconia*. This produces salts with both acids and bases, thus:—

Sodium zirconate..... Na_2ZrO_3.
Zirconium sulphate..... $Zr(SO_4)_2$.

The formation of the sodium salt may be represented as follows:—

$$\underset{\text{Zirconium oxide.}}{ZrO_2} + \underset{\text{Carbonate of sodium}}{Na_2CO_3} = \underset{\text{Sodium zirconiate.}}{Na_2ZrO_3} + \underset{\text{Carbon dioxide.}}{CO_2}.$$

The sulphate is formed according to the following equation:—

$$\underset{\text{Zirconium oxide.}}{ZrO_2.} + \underset{\text{Sulphuric acid.}}{2H_2SO_4} = \underset{\text{Zirconium sulphate.}}{Zr(SO_4)_2} + \underset{\text{Water.}}{2H_2O}.$$

THE ANTIMONY GROUP OF METALS.

ANTIMONY.	TANTALUM.	BISMUTH.
NIOBIUM.	VANADIUM,	

The individuals of this group form pentoxides (M_2O_5), which give rise to acids. In this respect the group closely resembles those of the group of non-metallic elements comprising nitrogen, arsenic and phosphorus. Only two, antimony and bismuth, are employed in medicine, in any form. The others are rare elements of no practical importance.

ANTIMONY. (*Stibium.*)

Symbol, Sb. Atomic weight,† 122.0.

Antimony is found in many minerals, but the chief ore is stibnite, a trisulphide of antimony, Sb_2S_3. To obtain the metal, the mineral is melted, when the sulphide of antimony in a fused state is produced. This is roasted in contact with air to convert it into oxide of antimony. The crude oxide of antimony is then mixed with powdered charcoal which has been saturated with a solution of crude tartar, and the mixture is roasted. Or, the purified sulphide of antimony is fused with metallic iron; sulphide of iron and metallic antimony results. The antimony is then run into moulds. Antimony is a silver-white metal, hard, brittle, and may be powdered in a mortar. It forms a crystalline mass, has the specific gravity of 6.7 or 6.8, and melts at 425° C. Cold air does not affect it, but it oxidizes when heated in the atmosphere, and volatilizes if heated in a current of hydrogen gas. Dilute sulphuric acid does not affect it, but both cold nitro-muriatic acid and hot hydrochloric acid dissolve it, and hot sulphuric acid produces with it sulphate of antimony. Phosphorus, arsenic, and members of both the chlorine and sulphur group of elements unite with it directly. Antimony is

† Recent experiments of Prof. Jos. P. Cooke give 120 as the atomic weight of antimony.

used in making some alloys, and a form of the metal produced by precipitation with zinc from solution of trichloride of antimony is known in the arts as *antimony black*, and is used to give to models and casts a surface which resembles steel. This must not be confounded with the ordinary *black antimony* of commerce, which is purified and powdered trisulphide of antimony.

Antimony and Potassium Tartrate.†— Formula, $C_4H_4K(SbO)O_6$. Molecular weight, 324.64.

Synonyms.—Antimonii et potassii tartras. Tartrate of antimony and potassium. Tartar emetic. Antimonium tartaratum.

This preparation has been used in medicine many years. It is made by dissolving oxide of antimony in solution of bitartrate of potassium, and permitting the crystals to separate. The following is essentially the process of the U. S. P. Mix two troy ounces of powdered oxide of antimony with two and one-half troy ounces of bitartrate of potassium, and mix the powder with sixteen fluid ounces of water and boil for an hour, supplying water to replace that which evaporates. Filter the solution while hot and place the filtrate in a cool situation to crystallize. Tartar emetic is always powdered before it is thrown upon the market. The crystals are colorless and transparent when recent, but become opaque upon exposure to the air from loss of water. They have the formula,

$$2(C_4H_4K(SbO)O_6).H_2O,$$

and impart at first a sweetish taste, followed by a strong metallic impression. The important impurity to be anticipated is cream of tartar, and this may be detected "by dissolving twenty-four grains of the salt in a fluid ounce of warm water, and cooling the solution to 10° C. (50° F.), when, if unadulterated, it will remain clear."

† Recent experiments by Prof. F. W. Clarke and H. Stallo, indicate that this compound is the potassium salt of a complex acid—$HSbC_4H_4O_7$. From this view the molecule of tartrate of antimony and potassium would be represented by the formula $KSbC_4H_4O_7$. (See *American Chemical Journal*, vol. ii. No. 5.)

Arsenic (National Dispensatory) is also liable to be present (from the oxide of antimony) if the solution of the salt has been evaporated very low before crystallization. This impurity is not likely to occur if care be used in the operation. The certain detection of small portions of arsenic is rather troublesome. (See page 155.) It is proposed to detect arsenic in tartar emetic by heating a few grains of the salt in a dry test tube until it blackens and gradually reaches redness. Arsenic will be present if a perceptible odor of garlic is evolved. Tartar emetic is a violent emetic, and in over doses, an irritating poison. The treatment for this and poisoning by other salts of antimony is given under antimony trichloride, or butter of antimony.

Antimony Trichloride.—Formula, $SbCl_3$. Molecular weight, 228.11.

Synonym.—Butter of antimony.

At one time butter of antimony was made by dissolving trisulphide of antimony in hydrochloric acid, evaporating the solution to small bulk, and then distilling the product. It was a white crystalline semi-transparent mass. At present the "butter of antimony" of commerce is a solution of trichloride of antimony, obtained by dissolving one part of trisulphide of antimony in four parts of hydrochloric acid by means of heat, and after filtration of the solution, evaporating it until it is reduced to two parts. It is impure from presence of iron and other metals, and has in consequence a red color. If distilled, a colorless solution of $SbCl_3$ results, but this, for use as a caustic, possesses in reality little advantage over the crude butter of antimony. Butter of antimony is a powerful and penetrating caustic, and is a violent poison. In cases of poisoning by this or other preparations of antimony, infusions of astringents, tannic acid in solution, or magnesia in milk, should be given. Fortunately *tartar emetic* causes violent vomiting, so that the use of emetics is unnecessary, but in case vomiting does not follow, the stomach pump must be used.

Antimony Trioxide. — Formula, Sb_2O_3. Molecular weight, 291.88.

Synonyms.—Antimonii oxidum. Oxide of antimony.

This occurs native as the mineral valentinite. It may be made artificially by dissolving sulphide of antimony in hydrochloric acid mixed with nitric acid, then diluting the solution with water and washing the precipitate until it ceases to be of acid reaction. The mass is then warmed with carbonate of sodium, and washed until the washings do not affect solution of nitrate of silver. The precipitate is then dried. The explanation of the process is as follows: Sulphide of antimony dissolves in hydrochloric acid to form antimony trichloride, $SbCl_3$, sulphide of hydrogen escaping, thus:—

$$\underset{\text{Antimony sulphide.}}{Sb_2S_3} + \underset{\text{Hydrochloric acid.}}{6HCl} = \underset{\text{Antimony trichloride.}}{2SbCl_3} + \underset{\text{Hydrogen sulphide.}}{3H_2S}.$$

The nitric acid present forms ferric chloride with the iron, which is separated by subsequent washing from the precipitated basic antimony chlorides. These are converted by washing with water into more basic oxy-chlorides, which ultimately, by the action of carbonate of sodium, form, by double decomposition, trioxide of antimony, and chloride of sodium. The preparation of this substance is very disagreeable, unless flues be provided to carry off the deleterious fumes, and few care to perform the operation for the sake of experiment. Oxide of antimony is a pale grayish-white powder, insoluble in water, but very soluble in warm hydrochloric acid, tartaric acid and caustic alkaline solution. It turns yellow when heated, but assumes its natural tint upon cooling. It fuses at a red heat, forming a deep yellow liquid, which upon cooling forms a semi-crystalline mass, of a buff color. Oxide of antimony is seldom employed excepting in the preparation of tartar emetic.

Other Oxides of Antimony.—There are two oxides of this element besides trioxide. These, *antimony tetroxide*, Sb_2O_4., and *antimony pentoxide*, Sb_2O_5., are of no use in medicine.

Antimony Trisulphide—Formula, Sb_2S_3. Molecular weight, 339.94.

Synonyms.—Antimonii sulphuretum. Sulphuret of antimony. Black antimony. Crude antimony.

The common mineral of antimony is this sulphide. We mentioned this fact when treating of the metal and its production, and described the manner in which the sulphide is prepared as a step in obtaining metallic antimony. If the crude fused sulphide is powdered we have the *black antimony* of commerce. It contains various metallic sulphides, such as iron, lead, copper and sulphide of arsenic, but these seem not to affect its use for general purposes. In preparing the salts of antimony for medicinal use such impurities are separated. Crude trisulphide of antimony occurs in masses of an ash gray color when broken and presents a metallic fracture. It yields a more or less black powder, and it is said that this as found in commerce is liable to adulteration with powdered coal. This can be detected by dissolving the sulphide in hot muriatic acid, when coal will remain. This compound may be made in a state of purity by passing a current of sulphide of hydrogen through solution of trichloride of antimony. An orange-red precipitate of hydrated trisulphide is produced, and this upon drying and heating becomes anhydrous, forming a deep black powder.

Antimonium Sulphuratum, or *sulphurated antimony*, is a mixture of trisulphide (Sb_2S_3) and pentasulphide (Sb_2S_5), with small amounts of trioxide (Sb_2O_3) of antimony. It is a reddish-brown powder, and is made by boiling together a mixture of antimony trisulphide, liquor potassæ, and distilled water, and adding to the solution sulphuric acid until it ceases to produce a precipitate. This precipitate is to be well washed with distilled water, and dried It seems strange that a mixture such as the above should, at this day, occupy a position in both the Pharmacopœia of the United States and of Great Britain; yet such is the case

Antimonii Oxysulphuretum, U. S. P., or *oxysulphuret of antimony,* has somewhat of a reputation under the name, *Kermes mineral.* This is a mixture of trisulphide of antimony (Sb_2S_3), and trioxide of antimony (Sb_2O_3). It is made by boiling trisulphide of antimony with a solution of carbonate of sodium, permitting the mixture to subside, and draining and washing the precipitate. This is a soft powder, of a more or less reddish-brown color. It is odorless and tasteless, and by means of a microscope is shown to be a mixture of amorphous antimony trisulphide and crystalline antimony trioxide.

Antimony Pentasulphide, Sb_2S_5.

This compound is known as *golden sulphuret of antimony.* It is a yellowish-red powder, soluble in solutions of the alkalies, and carbonates of the alkali metals. It is almost out of use in medicine.

In addition to the compounds of antimony we have mentioned there are many others which are of interest to the general chemist, but of no importance in medicine.

Tests for Antimony.—If a compound of antimony be heated on a carbon splinter a metallic bead results. Arsenic under like circumstances is volatilized. If a compound of antimony be introduced into the flask (A fig. 50) and the experiment be conducted according to the directions given for testing for arsenic (page 154), a metallic stain will result. This can be distinguished from arsenic in the manner directed upon page 155. The antimony film is soluble in chlorinated soda solution, NaOCl (Labarraque's solution), and sulphide of ammonium.

BISMUTH.

Symbol, Bi. Atomic weight, 210.0.

This metal is found native, and as the trioxide, Bi_2O_3, and, to a slight extent, as the trisulphide, Bi_2S_3. It is found sparingly in several minerals. Bismuth was formerly known under the name *marcasite.* The metal is obtained from the ore by processes unnecessary for us to consider, one being simply to heat the ore and run the impure molten metal into pots.

Crude metallic bismuth contains, as impurities, arsenic, copper, iron, silver, and often other metals. It may be purified according to the B. P. by melting it in contact with potassium nitrate and stirring the mixture well until a solidified slag forms upon the surface of the metal. The bismuth of commerce is usually contaminated with arsenic, and often with silver, both of which, manufacturers of the bismuth salts separate during the process of manipulation.

Properties.—Bismuth has a grayish-white color with a reddish tinge when viewed from certain positions, and it has a bright metallic lustre. It may be obtained in form of crystals by melting a pot of bismuth and breaking the crust which forms upon the surface as it cools, and then pouring out the liquid. Bismuth melts at 270° C. and has the specific gravity 9.823 (Holzmann). It unites directly with chlorine and the other elements of that group, and is readily dissolved by diluted nitric and nitro muriatic acids. It is insoluble in cold hydrochloric and sulphuric acids, but dissolves in hot sulphuric acid with formation of sulphur dioxide.

Bismuth forms valuable alloys, one, *fusible metal*, composed of bismuth, lead, tin, and cadmium, being used in the arts for stereotyping, and for making models. This metal fuses at a temperature less than the boiling point of water, and has the important property of expanding during the act of solidification, thus filling the mold perfectly.

Bismuth forms four oxides:—

Bismuth dioxide..... Bi_2O_2.
Bismuth trioxide..... Bi_2O_3.
Bismuth tetroxide ..... Bi_2O_4.
Bismuth pentoxide..... Bi_2O_5.

None of the above are used in medicine.

Bismuth trihydroxide, $Bi(OH)_3$, is easily made by precipitating a solution of trinitrate of bismuth with solution of caustic potash, thus:—

$$\underset{\text{Bismuth trinitrate.}}{Bi(NO_3)_3} + \underset{\text{Caustic potash.}}{3KOH} = \underset{\text{Bismuth trihydroxide.}}{Bi(OH)_3} + \underset{\text{Potassium nitrate.}}{3KNO_3}.$$

This is not used in medicine, but with acids forms salts, thus:—

$$\underset{\text{Bismuth trihydroxide.}}{Bi(OH)_3} + \underset{\text{Nitric acid.}}{3HNO_3} = \underset{\text{Trinitrate of bismuth.}}{Bi(NO_3)_3} + \underset{\text{Water.}}{3H_2O},$$

It dissolves in solution of citrate of ammonium to form *ammonio-citrate of bismuth.*

Bismuth Carbonate. — Formula, $2(BiO)_2CO_3.H_2O$. Molecular weight, 1041.50.

Synonyms —Bismuthi subcarbonas. Subcarbonate of bismuth. Basic carbonate of bismuth. Carbonate of bismuth.

This compound may be made by pouring a solution of trinitrate of bismuth into a solution of carbonate of ammonium, when the basic bismuth carbonate precipitates. According to the Pharmacopœia of the United States it is made by a round-about process, designed to separate arsenic from crude bismuth, carbonate of sodium being finally employed as a precipitant.

Subcarbonate of bismuth is usually of a yellowish-white color, but when freshly precipitated it is white. Sometimes it is very light and bulky as found in market, and again it is heavy and compact. This variation results in accordance with the temperature of the solutions employed during precipitation.

Subcarbonate of bismuth dissolves in nitric acid, and if lead and chlorine are absent will not yield precipitates upon the addition of dilute sulphuric acid or solution of nitrate of silver. Arsenic may be detected by Marsh's test, or by placing a little of the subcarbonate of bismuth in a test-tube with some liquor potassæ, and plugging the mouth of the tube with a piece of cotton moistened with solution of nitrate of silver. Upon boiling the liquid, the cotton will turn black or gray if arsenic is present. Silver is shown by the action of light upon that portion of the subcarbonate next the bottle, turning it a gray color. Sulphide of hydrogen will affect bismuth subcarbonate in this manner, turning it gray, therefore the absence of this article must be insured.

Bismuth Citrate.—Formula, $BiC_6H_5O_7+Aq$.

Dissolve ten parts of subcarbonate of bismuth in fifteen parts of nitric acid; carbon dioxide escapes and trinitrate of bismuth results. Dilute this solution with fifteen parts of water, and dissolve in it nine parts of citric acid. Add to this, with constant stirring, a solution of caustic potash, nine and one-half parts in eighty parts of water; a white precipitate of hydrated citrate of bismuth results.† This can be freed from nitrate of potassium by washing. If dried it forms a white powder. This preparation is not employed in medicine, but if the moist and well washed precipitate be dissolved by the addition of a little ammonia water, a clear solution results which may be obtained in scales after the manner of making ammonio-citrate of iron. This is known as—

Bismuth Ammonio-citrate, and is employed extensively in this country, being, in fact, the only solution of a bismuth salt employed in medicine. The equation expressing the reactions which take place in making citrate of bismuth may be expressed as follows:—

$$\underset{\text{Bismuth subcarbonate.}}{(BiO)_2CO_3} + \underset{\text{Nitric acid.}}{6HNO_3} = \underset{\text{Bismuth trinitrate.}}{2Bi(NO_3)_3} + \underset{\text{Carbonic acid.}}{H_2CO_3} + \underset{\text{Water.}}{2H_2O.}$$

This explains the part of the reaction which results between the subcarbonate of bismuth and the nitric acid. The addition of water and citric acid produces no change until the potassium hydroxide is added, when the following results:—

$$\underset{\text{Bismuth tri-nitrate.}}{Bi(NO_3)_3} + \underset{\text{Citric acid.}}{H_3C_6H_5O_7} + \underset{\text{Caustic potash.}}{3KOH} + \underset{\text{Ammonium hydrate.}}{3NH_4OH} =$$

$$\underset{\text{Bismuth ammonio-citrate.}}{Bi(OH)_3\,(NH_4)_3C_6H_5O_7} + \underset{\text{Potassium nitrate}}{3KNO_3} + \underset{\text{Water.}}{3H_2O.}$$

† Some discussion has resulted from the study of this compound and ammonio-citrate of bismuth. Authorities are by no means agreed as to the chemical reactions which take place, or the constitution of the double salt which results. We prefer to consider that citrate of bismuth has the formula $BiC_6H_5O_7$, and that ammonio-citrate of bismuth is represented by $Bi(OH)_3$-$(NH_4)_3C_6H_5O_7+Aq$. thus in reality being a basic citrate of bismuth, or, if some so choose to consider, a solution of tri-hydroxide of bismuth in citrate of ammonium.

The British Pharmacopœia prepares an impure ammonio citrate of bismuth, and recognizes it as *liquor bismuthi et ammoniæ citratis*. It is contaminated, however, with potassium nitrate, and such as desire a solution of a bismuth citrate may well prefer the preparation made as we have indicated. This is sold on the market under the name *liquor bismuthi*.

Bismuth Trinitrate.—Formula, $Bi(NO_3)_3$. Molecular weight, 395.67.

Synonyms. — Nitrate of bismuth. Acid nitrate of bismuth.

This salt is of no use in medicine. It is made by dissolving metallic bismuth in dilute nitric acid. Crystals of the composition, $Bi(NO_3)_3.3H_2O$, are obtained upon evaporation of the solution and subsequent crystallization. These crystals are colorless and transparent, deliquesce in the air, and by heat are decomposed, finally leaving trioxide of bismuth. If a solution of trinitrate of bismuth be poured into *cold water*, a basic nitrate (oxynitrate) is produced as an insoluble precipitate. This is used extensively in medicine under the name—

Bismuth Sub nitrate.

It has the composition $Bi(OH)_2NO_3$, or $BiONO_3.H_2O$, as some prefer. This is a heavy crystalline white powder, which is odorless, and not of an acid taste. It reddens moistened blue litmus paper, and dissolves perfectly in nitric acid. It should not show the presence of arsenic by the test we have given under bismuth sub-carbonate (page 339), although much of the sub-nitrate of bismuth of the market is contaminated with traces of that element.† Silver will sometimes be present in sub-nitrate of bismuth in amounts sufficient to produce a gray color upon exposure of the salt to light.

† Admitting this point the writer is still of the opinion that the expressed fears of poisoning from this source are much overdrawn, although occasionally it is present in considerable proportion. I am informed by Mr. Charles Rice that he has a sample of subnitrate of bismuth that contains one per cent. of arsenic.

Bismuth Oxy-chloride, BiOCl, or $BiCl_3Bi_2O_3.H_2O$, is a white powder employed as a cosmetic. It may be made by pouring a solution of trinitrate of bismuth into a solution of chloride of sodium, or by dissolving the *trichloride* of bismuth in an excess of water. The other bismuth compounds are not of therapeutical interest, although a *tannate* and a *valerianate* of bismuth have been mentioned.

Tests for Bismuth.—If sulphide of hydrogen be passed through the solution of a bismuth salt, a black-brown sulphide results. This refuses to dissolve in ammonium sulphide but dissolves in nitric acid. If solution of potassium chromate be added to the solution of a bismuth salt, a yellow precipitate results. This dissolves in dilute nitric acid, and is insoluble in solution of caustic potash.

TANTALUM AND NIOBIUM.

These two elements are generally found associated, both are rare and neither is of any practical use.

TANTALUM.

Symbol, Ta. Atomic weight, 182.0.

Tantalum was discovered in 1803, by Ekeberg; it has never been obtained pure, although Berzelius produced a black powder capable of assuming a metallic lustre. This burned to an oxide when heated in the air, dissolved in hydrofluoric acid, hydrogen being evolved. It united with sulphur vapor and chlorine gas when heated in either, burning brightly. Tantalum forms two oxides, *tantalum tetroxide,* Ta_2O_4 and *tantalum pentoxide,* Ta_2O_5.

NIOBIUM.

Symbol, Nb. Atomic weight. 94.0.

It was discovered by Hatchett (1801), in an American mineral, Columbite, and named by him *Columbium.* It was afterward assumed (Wollaston) to be identical with tantalum, but Rose, in 1846, positively showed that the two elements existed, giving the name *Niobium* to that

which had nearly half a century previously received the name Columbium. Niobium (Columbium) was obtained in a film by Bomstrand, and as a steel-gray crust by Roscoe. It is soluble in sulphuric acid, but not in the other mineral acids It forms three oxides, and a hydroxide. The salts of Niobium are of no use.

VANADIUM.

Symbol, V. Atomic weight, 51.2.

In 1801 Del Rio discovered this element in an ore from Mexico, *vanadate of lead.* He named the new element Erythronium from its red color, but his conclusions were distrusted and considered erroneous. Sefstrom again discovered it in 1830, and gave it the name *Vanadium.* It is made pure by reduction of the dichloride in hydrogen gas, a difficult and tedious experiment. Vanadium is a whitish gray colored powder, appearing under the microscope as a brilliant metallic crystalline mass. It oxidizes very slowly in the air, but burns brightly in a flame, or when it is heated in oxygen gas. Its specific gravity is 5.5. It is violently oxidized by nitric acid, dissolves in hydrofluoric acid, but neither cold hydrochloric nor sulphuric acid, dilute or concentrated, affects it. Vanadium forms five oxides, all of which produce salts. It is unnecessary for us to describe these, as neither the metal nor a single salt of it is of any medicinal use.‡ One peculiarity of vanadium is its property of uniting directly with *nitrogen.* The affinity between these substances is such that oxide of vanadium gives up its oxygen, and the metal unites with nitrogen, when the oxide is heated in an atmosphere of mixed nitrogen and hydrogen.

‡ The salts of Vanadium are now used in dyeing, and for making black ink, and for dyeing. The ore is supplied from Swedish mines.

THE GOLD GROUP OF METALS.

GOLD.	PALLADIUM.	PLATINUM.
IRIDIUM.	RHODIUM.	RUTHENIUM.
OSMIUM.		

Gold and Platinum refuse to unite directly with oxygen, even when heated in that gas.

Palladium, Rhodium and Iridium unite with oxygen when gently heated in the air, but these oxides decompose when the heat is increased, oxygen escapes and the element remains.

Ruthenium and Osmium unite with oxygen and give rise to volatile oxides which do not decompose at any temperature.

The lower oxides of the foregoing group act as weak bases, the higher oxides are acid producers.

Gold is the only metal of this group of elements that is used in medicine.

GOLD. (*Aurum.*)

Symbol, Au. Atomic weight, 196.2.

Gold was named from its lustre or shining color. It is generally found native, alloyed more or less with silver. It is also found in some minerals in combination, and is present in the waters of the ocean. It is very widely distributed, being found in almost every country of the globe. The processes for its extraction from sand and quartz are well understood, and do not require mention by us. Gold has been considered a precious metal since the earliest times, and is the only element having a bright yellow color. It crystallizes in the regular system, and when small, natural crystals are found connected together, "in hair-like filaments," it is termed *moss gold.* Gold is the most ductile of metals and to such an extent can it be hammered, that 280,000 thicknesses of gold leaf will only occupy one inch in thickness, and one grain of gold will gild two miles of silver wire. Gold may be so finely divided by precipita-

tion that, notwithstanding the specific gravity of the metal, it remains suspended in water, imparting to the mixture a blue color by transmitted light, and a purple-red color by reflected light. Gold leaf permits the passage of light, and in this case the green ray is transmitted. Gold has the specific gravity 19.265 (Matthiessen), and melts at 1381° C. (Pouillet). It remains unaltered by exposure to water or oxygen of any temperature and for any length of time. Pure gold is very soft, and for coinage and for jewelry requires to be alloyed with other metals. For this purpose both copper and silver are employed; if copper, a red color is imparted to the alloy, the silver alloy being light yellow in color.

Gold trinkets are sold according to a certain percentage of gold, the standard of which is called "carat." Thus, 24 carat is *pure gold*, and 22, 18, 15, 12 carat, etc., signify such proportions of pure gold in twenty-four parts of the alloy. Gold forms an amalgam with mercury, and occasionally is found native as an amalgam. Gold is not affected by even the concentrated mineral acids, but a mixture of muriatic and nitric acids dissolves it freely, nascent chlorine being the agent, and which combines with the gold to form *chloride of gold*, or gold trichloride, $AuCl_3$. There are two known oxides of gold, neither produced directly from the metal.

Gold Monoxide, or *aurous oxide*, Au_2O, may be made by decomposing aurous chloride (AuCl) with solution of caustic potash. *Gold trioxide* or *auric oxide*, Au_2O_3, is obtained by heating the hydroxide to 100° C. (212° F.)

Gold Hydroxide, or *trihydroxide of gold*, $Au(OH)_3$, is made by heating a solution of trichloride of gold ($AuCl_3$) with calcined magnesia, and afterward washing the precipitate well with nitric acid.

Fulminating Gold.—This is *ammoniacal auric oxide*, $Au_2O_3(NH_3)_4$. It is made by acting upon gold hydroxide with ammonia. There is a simpler process for making this compound, but none should make it for experiment, and we do not care to give it. Fulminating gold is among the most dangerous of explosives, a jar being sufficient to

cause its decomposition. Gold gives rise to two classes of salts, *aurous* salts and *auric* salts. These follow the general rule for such compounds, the proportion of gold in the aurous salts being greater than that of the auric salts, thus:—

{ Aurous chloride.....AuCl.
{ Auric chloride.....$AuCl_3$.

{ Aurous bromide.....AuBr.
{ Auric bromide.....$AuBr_3$.

One salt of gold is used in medicine, the—

Auric Trichloride, or *chloride of gold,* $AuCl_3$.—To prepare it, dissolve gold coin it nitro-muriatic acid. Evaporate the solution to dryness, dissolve the residuum in water, and pour the solution into an excess of solution of sulphate of iron (ferrous sulphate). A black precipitate of metallic gold is produced. This must be well washed by decantation, dissolved in nitro-muriatic acid, and the solution evaporated to dryness. The residue is now to be dissolved in a little water and filtered, and the filtrate evaporated to dryness, when the salt must be immediately placed in well stoppered vials. Chloride of gold has a dark red color in mass; if powdered it is yellowish red. There should be no greenish tinge, and the salt should dissolve perfectly in alcohol, ether and water.

Chloride of Gold and Sodium.

There is a double salt of chloride of gold and chloride of sodium. It has the composition $NaAuCl_4 2H_2O$, and is made by adding chloride of sodium in proper proportion to solution of chloride of gold and afterward evaporating the solution to crystallization. The crystals are rhombic prisms, permanent, and of a yellowish red color, and are the chloride of gold and sodium of the French Codex. In this country, however, *chloride of gold and sodium* is made by dissolving equal weights of chloride of gold and chloride of sodium in water, and evaporating the solution to dryness. This has an orange-yellow color and is the form of gold mostly employed in medicine. The other salts of gold are unimportant from a therapeutical point of view.

Tests for Gold.—Add solution of sulphate of iron to the solution of any salt of gold, a black precipitate is produced. Rub this between two smooth metallic surfaces, a bright yellow metallic luster results. Add the precipitate to concentrated nitric acid, it will not dissolve. Add now hydrochloric acid, solution results.

2. Add solution of ferric chloride to solution of stannous chloride ($SnCl_2$) until a permanent yellow color is produced. Dilute this solution. Then place the solution to be tested for gold in a beaker glass upon a sheet of white paper. Dip a glass rod into the mixed iron and tin solution, then draw it through the liquid within the beaker. If a trace of gold is present, a blue or purple streak will follow the glass rod. (Galloway.) This test is upon the principle that stannous chloride produces the "purple of Cassius" with gold solutions. If a drop of the solution of a salt of gold be placed upon a sheet of paper, and a drop of solution of stannous chloride be added, a purple-red color results.

PLATINUM.

Symbol Pt. Atomic weight, 196.7.

Platinum has been known since the middle of the sixteenth century, but it did not attract general attention until two centuries later. Platinum is found in small amounts in most countries, but a comparatively few localities furnish the platinum of commerce. Platinum is only found in the native state, but this is usually impure, being contaminated with all of the elements of this group, and iron, copper, and other substances. It is occasionally found crystallized as cubes, but usually in the form of flattened or irregularly shaped grains. Pure platinum has a white color, the specific gravity 21.5 (Deville and Debray), and is soft as copper. It can be readily melted at a white heat. It is (excepting gold and silver) the most malleable of all metals. Platinum refuses to melt at the greatest temperature of a blast furnace, if in large mass. A very fine wire fuses in the flame of a Bunsen jet, and considerable amounts can be melted in the oxy-hydrogen flame. Heated platinum permits the passage of hydro-

gen gas through it with comparative rapidity. This is explained upon the principle that the surface next to the gas absorbs it readily, and the outer surface throws it off. The vapors of many inflammable liquids, and certain gases, will cause a thin sheet of platinum to become red-hot, often attended by ignition or even explosion. This can be shown by holding a clean, warm, piece of platinum foil in a stream of *dry* hydrogen gas as it escapes into the atmosphere. The foil quickly becomes red-hot and the hydrogen is ignited.

Spongy platinum is a form of metallic platinum which presents a very large surface. It is made by igniting the double chloride of platinum and ammonium. It has the power of condensing large quantities of oxygen. *Platinum black* is another form of platinum. It is a black powder, and when carefully prepared in a vacuum absorbs oxygen so rapidly upon exposure to the air as to become red hot. Platinum is valuable to the chemist. It is not affected by oxygen, water, and sulphuric or nitric acid, at any temperature, therefore platinum crucibles and retorts are used for many experiments where other metals would be useless. Many alkalies, alkaline cyanides, burning charcoal, phosphorus, arsenic, (sulphur if alkali is present), and metals, attack platinum crucibles. Care must be taken, therefore, to guard against such bodies. Platinum should retain its brightness, and should not lose weight in either boiling nitric or hydrochloric acid.

Platinum dissolves readily in nitro-muriatic acid, *chloroplatinic acid*, H_2PtCl_6 being formed. This is usually called *chloride of platinum.* The hydrogen of this acid replaced by different metals forms salts, and it combines with many alkaloids to form insoluble compounds. For this reason the acid (so-called chloride of platinum) is of use in organic analysis as an alkaloidal reagent. In all cases the absence of ammonia must be ensured, as it forms *ammonium platinichloride* with chloroplatinic acid. This is a very insoluble salt of the composition $(NH_4)_2PtCl_6$, and which may readily be taken for the precipitate of an alkaloid.

There are two classes of salts of platinum, *platinous* salts and *platinic* salts, of which the following are examples:—

{ Platinous chloride, or platinum dichloride..$PtCl_2$.
{ Platinic chloride, or platinum tetrachloride $PtCl_4 5H_2O$.

{ Platinous iodide.. ....PtI_2.
{ Platinic iodide.....PtI_4.

In addition there are many and complex compounds of this element with other substances, of great interest to students of pure chemistry but of no practical value.

PALLADIUM.

Symbol, Pd. Atomic weight, 106.2.

This element was discovered in 1803, by Wollaston, and named from the planet *Pallas*, which was discovered during the preceding year. It is found associated with most platinum ores, and is often found alloyed with gold. It is usually obtained as a by-product during the purification of platinum, but the metal is occasionally found tolerably pure (usually containing iridium and platinum) in the form of grains like platinum, which we have described. Palladium has a white color, specific gravity 11.4 (Deville and Debray), and melts more easily than any other metal in the gold group of elements. It assumes upon the surface a dark blue or a violet color when it is heated to dull redness, but at high temperatures it regains its brightness. Hot nitric acid, and hydrochloric acid into which chlorine is passed, dissolve it. Palladium has been used in the arts to a slight extent. It is unalterable in the air and has, therefore, been employed to coat astronomical instruments. It does not alter when it is exposed to sulphide of hydrogen, therefore dentists have employed it instead of gold. None of the compounds of palladium are of practical interest.

RHODIUM.

Symbol, Rh. Atomic weight, 104.1

Wollaston discovered this element, and named it in reference to the rose red color of its salts. It usually occurs in platinum ore, and is obtained from the solution which remains after precipitation of ammonium platinic chloride. Rhodium has the specific gravity 12.1 (Deville and Debray), and the color and appoarance of aluminium. It is one of the most difficult of elements to fuse, platinum melting at a lower temperature.. Rhodium, if pure, is almost insoluble in acids, not excepting nitro-muriatic acid, but if it be alloyed with many other elements it dissolves in the latter named compound. Rhodium is rare, and its salts are of no value in the arts or in medicine.

IRIDIUM.

Symbol, Ir. Atomic weight (Tennant), 196.7.

This element was discovered by Tennant, in 1803 and 1804, in the residue left from solution of platinum ore, in which it is found as an alloy. Iridium has the specific gravity of 22.38, and when polished, resembles polished steel. It is slightly malleable at a white heat, but is very brittle when it is cold. Spongy iridium may be prepared after the manner for making spongy platinum; this will oxidize in the air, but the metal in mass will not oxidize at all. Pure iridium will unite with chlorine at dull redness, but nitro-muriatic acid will not affect it, unless the metal be alloyed with platinum, in which case this acid dissolves it. When alloyed with platinum, (iridium one part, and platinum nine parts,) it forms the *standard-metal* which is used to make standard measures. This alloy is very hard, very elastic, and perfectly unchangeable in the air. It takes a high polish, and melts at a higher degree of temperature than is required to fuse platinum. The compounds of iridium are of no interest to physicians or pharmacists.

RUTHENIUM.

Symbol, Ru. Atomic weight, 103 5.

This element is found in platinum ore. It was first positively identified by Claus (1845), although Osann (1828) stated that platinum ores from the Ural mountains contained *three* new metals. One of these he named *ruthenium*, and as Mr. Claus actually found it in small quantity in the substance Osann had prepared for ruthenium oxide, he (Claus) accepted the name. The claims for the other two elements Osann stated to have discovered, were shown to be unfounded. Ruthenium has the specific gravity 12.261 at 0° C. (32° F.) It is hard, brittle, and excepting, *osmium*, the most difficult to fuse of the metals of the gold group, likewise, excepting osmium, it combines more readily with oxygen than any other metal of the group. The pure metal combines with chlorine at a red heat, but nitro-muriatic acid scarcely affects it. Ruthenium is very rare, and is never employed in the arts or in medicine either in elementary form or as a compound.

OSMIUM.

Symbol, Os. Atomic weight, 198.6

Osmium was discovered by Tennant at the time he discovered iridium (1804), and in the same ore. He named it *osmium* after the peculiar odor of its oxide, which is volatile. It is readily separated from the other members of the gold group by taking advantage of this volatile nature of the tetroxide (OsO_4), and from this compound Deville and Debray obtained the element.

Osmium is the heaviest known body, having the specific gravity of 22.477 (Deville and Debray), and it has *never* been melted. For these properties it is remarkable among metals, and also for the fact that to volatilize even minute amounts of it, which can only be accomplished by means of the oxy-hydrogen blow-pipe, is an extremely dangerous operation.

Osmium tetroxide, OsO_4, or *osmic acid,* is formed by several processes. It possesses the most penetrating properties, is the most irritating of any known substance, and experimenters are required to use the utmost caution that it does not escape into the atmosphere. Deville nearly lost his sight from its action, after a very slight exposure to its vapor, and a very small amount of the vapor, if mixed with the air, will give rise to the most serious inflammation of mucous surfaces.

Osmiridium is the name given to an alloy of osmium and iridium, and which is used for the points of gold pens. This alloy will not oxidize, is not attacked by acids, and is not magnetic. In consequence of the latter property it is used for the bearings of compasses. None of the compounds of osmium are of interest to pharmacists.

We have now considered the last known element of the gold group, and all of the known metals. We have endeavored to notice fairly each of the elements, and such of their compounds as are used in pharmacy or in medicine, excepting the carbon compounds included under the section to follow.

ORGANIC CHEMISTRY.

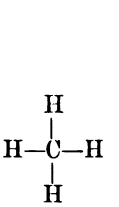
Under this head we shall study such important constituents of the vegetable and animal kingdoms, and their derivations, as are used in medicine. At one time it was supposed that all organic bodies were produced by what is called *vital force*, and could not be made artificially. Since that day many products of the vegetable and of the animal kingdoms have been artificially prepared, and the old line of demarkation between organic and inorganic chemistry is abandoned. At present the term organic chemistry is applied to the study of carbon compounds, and thus we may consider that the department to follow is a continuation of the study of our last non-element, carbon.

In this work, a fragment only of organic chemistry can be considered, as the majority of the compounds are of no interest in medicine We shall not, therefore, endeavor to closely follow the advanced views of theorists, as the matter at our command will not permit a systematic course, and without such it is useless to endeavor to grasp the subject. Such as are generally interested in this important section of chemical science, are referred to works devoted to organic compounds. Of first importance is the combination of carbon and hydrogen, known as the hydrocarbons.

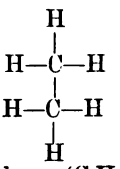
SATURATED HYDROCARBONS.

The atom of carbon (C^{IV}) can combine with four atoms of hydrogen. If it is in combination with this number, it is saturated and can not combine with another. If we have two carbon atoms combined in one molecule, each

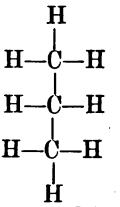
saturates one bond of attraction of the other, therefore, only *two* additional hydrogen atoms can be taken up. If now another carbon atom be introduced into the compound, it will be necessary to also introduce two hydrogen atoms. Each of the aforenamed molecules will be *saturated*, and in each the increase or decrease will be by the addition or subtraction of CH_2. Such a series can be carried theoretically to infinity. The series is called *homologous*, in consequence of each compound differing from the one below or above by abstraction or addition of CH_2. As examples we give the following graphic formulæ:—



Methane (CH_4).



Ethane (C_2H_6).



Propane (C_3H_8)

Unsaturated Hydrocarbon Radicals

Saturated hydrocarbons part with *one* atom of hydrogen, the result being radicals known as *alcohol radicals*. These radicals can not exist free, but they are basic in action and behave like metals. They displace the basylous hydrogen from acids, and uniting with the acidulous radicals form salts which are permanent. Thus from the saturated hydrocarbon *ethane* C_2H_6, we have C_2H_5 by abstraction of one atom of hydrogen, which is theoretically the radical of ordinary alcohol, alcohol being considered hydroxide of ethyl, C_2H_5OH, (See alcohol.) When the unsaturated alcohol radical unites with an acid radical, oxygen, or with a haloid element, an ether is formed, (see ethers).

ALCOHOLS.

There are a number of substances which the chemist of the present accepts as alcohols, and which bear little resemblance to ordinary alcohol. The term *alcohol* is

now applied to a large number of organic compounds which seem to contain hydroxyl united with unsaturated hydrocarbon radicals (basylous). These radicals are sometimes called *alcohol-radicals*, and have various combining powers, seeming to unite with one, two, three or more molecules of hydroxyl, thus :—

Alcohol radicals.	Hydroxyl.		Alcohols
Ethyl.. ....C_2H_5	+ OH	= C_2H_5OH	.....Ethyl alcohol.
Ethylene...C_2H_4	+ $(OH)_2$	= $C_2H_4(OH)_2$	..Ethylene alcohol.
Glyceryl....C_3H_5	+ $(OH)_3$	= $C_3H_5(OH)_3$	..Glyceryl alcohol.

A large number of alcohols are known and recognized by theoretical chemists, but few, however, are of use in medicine. Many of these (glycerin for example) are not classed by ordinary pharmacists and physicians with the alcohols.

Ethyl Alcohol.—Formula, C_2H_5OH.

Synonyms.—Ethyl hydrate. Ethylic alcohol.

Alcohol results from fermentation of saccharine matters through the agency of yeast. In practical work we believe that the starch of grain is first changed into corn-sugar (glucose), as follows :—

$$\underset{\text{Starch.}}{3C_6H_{10}O_5} + \underset{\text{Water.}}{H_2O} = \underset{\text{Dextrin.}}{2C_6H_{10}O_5} + \underset{\text{Glucose.}}{C_6H_{12}O_6}.$$

The sugar or glucose then undergoes, under the influence of yeast the change—

$$\underset{\text{Glucose}}{C_6H_{12}O_6} = \underset{\text{Alcohol.}}{2C_2H_5OH} + \underset{\text{Carbon dioxide.}}{2CO_2}.$$

In addition to alcohol, several incidental products, such as fusil oil, result.

Pure (Absolute) Alcohol, can not be made by ordinary distillation, and to produce it the officinal spirit must be distilled from freshly burned lime, carbonate of potassium or calcium chloride. Pure (absolute) alcohol has the specific gravity of 0.794 at 60° F., a sharp, burning taste and a pleasant odor. It dissolves iodine, bromine, many salts of metals, most alkaloids, essential oils, resins, etc. When it is mixed with 50 per cent. of water—

Proof Spirit results. This has the specific gravity 0.936 at 60° F. (15.5° C.) and is the standard of proof adopted by the U. S. Revenue service. The following *proofs* are recognized.

Specific gravity, 0.931; per cent. alcohol, 52½; second proof.
Specific gravity, 0.925; per cent. alcohol, 55½; third proof.
Specific gravity, 0.920; per cent. alcohol, 58; fourth proof.

Stronger Alcohol, U. S. P.—This has the specific gravity 0.817.

Alcohol, U. S. P.

Synonyms—Spiritus rectificatis. Refined spirit. Spirit of wine.

This has the specific gravity 0.835, and contains eighty-five per cent., by weight, of absolute alcohol. It has a peculiar odor, but should not smell of fusil oil when diluted with water. Alcohol will mix with castor oil in all proportions, the mixture being transparent.

Diluted Alcohol is a mixture of officinal alcohol (s. g. 0.835) with an equal bulk of water. It has the specific gravity 0.941.

When potassium or sodium is added to pure alcohol, hydrogen gas is evolved, and *ethylate* of potassium, or of sodium, is formed. When pure alcohol is mixed with an excess of officinal nitric acid, oxidation and a violent reaction ensue. If nitric acid be added in proper amount to ordinary alcohol, and the mixture be distilled, nitrous ether (nitrite of ethyl) and aldehyd result. If sulphuric acid and metallic copper be added to the mixture of nitric acid and alcohol, the production of aldehyd is much decreased (see spirit of nitrous ether). Sulphuric acid abstracts the elements of water from alcohol, the result being so-called sulphuric ether (see *ether*). If one part, by measure, of alcohol be mixed with one part, by measure, of water, contraction results, the mixture decreases in bulk and increases in temperature. The admixture is immediately followed by a copious separation of air bubbles, which imparts a milky appearance that disappears quickly, however, as the bubbles of air rise and escape. Alcohol is extensively employed in the arts and in medicine.

Methylic Alcohol —Formula, CH_3OH.

Synonyms—Wood naptha. Wood spirit. Wood alcohol. Pyroligneous spirit. Methyl hydrate.

This substance is practically produced by the destructive distillation of wood. It is, as first obtained, very impure, from the presence of acetic acid and tarry products. After several redistillations, and filtration through charcoal, and redistillation from lime, it becomes colorless and constitutes the *wood alcohol* of commerce.

Wintergreen oil is mostly *methyl-salicylic ether*, and yields methylic alcohol and salicylate of potassium when decomposed with caustic potash. Methylic alcohol may also be produced synthetically from methyl hydride (CH_4).

Wood alcohol of commerce is unfit for medicinal use, and is usually very impure When pure it is colorless, of specific gravity 0.814, and burns with a blue flame. It has a peculiar odor which is very disagreeable to many persons, and its vapor irritates and inflames the eyes. In England alcohol mixed with ten per cent. of *methylic alcohol* is made for pharmaceutical purposes, and is free from government tax. In this country the law will not permit such a course, our Pharmacopœia does not recognize (wisely, we think) wood alcohol, and if any use it in the preparation of medicines it is an adulterant. In the arts, however, wood alcohol is extensively employed, as in varnish making, etc.

Amylic Alcohol.—Formula, $C_5H_{11}OH$,

Synonyms.—Fusil oil. Amyl hydrate. Oil of corn. Isopental alcohol. Fuselol.

This is a side product of the fermentation of grain and potatoes. In this country fusil oil is produced in large amounts by the Western refining distilleries, and it is used extensively by manufacturers of artificial fruit and liquor flavors, as well as in the preparation of valerianic acid. Fusil oil is colorless, of specific gravity 0.818, and boils at 132° C. (about 270° F.) It has a penetrating, disagreeable odor; a burning taste; dissolves some alkaloids, and is soluble in alcohol and ether. As found in commerce

it is impure from the presence of ordinary alcohol, which is best separated by repeatedly shaking the crude fusil oil with fresh portions of saturated solution of sodium chloride, then distilling it from calcium chloride. Fusil oil is poisonous, and the injurious effects of common whisky (outside of the alcohol) are due to this article, which to a great extent is separated by the refining process.

Glycerin.—Formula, C_3H_5OH.

Synonyms—Glyceryl hydrate. Glyceric alcohol. Propenyl alcohol.

This is a sweet, viscid substance, and is obtained as a by-product of the candle factories, resulting from the decomposition of tallow. In the preparation of lead plaster, glycerin is a result; and at one time it was obtained in this manner.

Neutral fats and oils are, in fact, the ethers of propenyl alcohol (glycerin) in which hydrogen is replaced by the fatty acids. Thus, three molecules of stearic acid ($C_{18}H_{35}O_2$) may be said to be combined with glyceryl, C_3H_5, to form the fat known as *stearine*, $C_3H_5(C_{18}H_{35}O_2)_3$. Glycerin is colorless and odorless, neutral to litmus paper, of specific gravity 1.26 when anhydrous, but in commerce is usually 1.25. It is soluble in water, alcohol, and chloroform, but not in ether. It should not turn brown if mixed with two measures of cold sulphuric acid, and "when diluted with water, it affords no precipitate with hydrosulphate of ammonium, ferrocyanide of potassium, nitrate of barium, oxalate of ammonium, or nitrate of silver." U. S. P.

When glycerin is heated over direct heat, it decomposes, evolving irritating fumes; it distills, however, in a current of superheated steam, and in commerce is purified in this manner. It is remarkable as a solvent, dissolving many salts, most vegetable tannates, tannin, and many alkaloids. Freshly precipitated tannate of quinine dissolves in glycerin to the extent of five grains to the fluid ounce, and the solution is scarcely bitter. When glycerin is heated with strong nitric acid, *nitro*-glycerin, $C_3H_5(NO_2)_3O_3$ results, one of the most dangerous of ex-

plosives. *Dynamite* is a mixture of nitro-glycerin and sand. Glycerin is extensively employed in medicine and in the arts. Our Pharmacopœia recognizes a class of solutions of medicines in glycerin, under the name "glycerita."

Mannite.—Formula. $C_6H_8.6OH$.

Synonyms.—Mannityl hydrate. Mannitic alcohol.

Manna is a white, concrete substance obtained from the incised trunks of certain trees, from which it exudes as a sticky, gummy, mass. It has a peculiar odor, and a sweet taste, containing sugar, extractive matter and *mannite*. Mannite is easily obtained by boiling manna in alcohol of the specific gravity 0.914, filtering the solution, and permitting it to crystallize. The crystals are purified by re-solution and crystallization from alcohol. Mannite is insoluble in ether; sweetish to the taste, and does not respond to the potassium hydrate and potassium tartrate test for glucose. It does not readily undergo the vinous fermentation; by the action of ordinary nitric acid is first converted into *saccharic* acid, and lastly into *oxalic* acid. Manna is used to a limited extent in medicine, but *mannite* is of use neither in the arts nor in medicine, excepting as a constituent of manna, where it exists to the extent of ninety per cent.

Carbolic Acid, C_6H_5OH, is an alcohol of the aromatic series, and is known to chemists as *phenyl hydrate*. In commerce it is called phenic acid or carbolic acid; we have thought best to describe it among the acids proper (page 91). It will not decompose the carbonates, and thus may be said to differ from true acids, excepting hydrocyanic acid (HCN). Carbolic acid is sometimes called "coal tar creasote," and in an impure form has been mixed with or sold for true—

Creasote.—This is a mixed product of the dry distillation of wood. It is oily; colorless; of specific gravity 1.046, U. S. P. (1.071 Br. P.) and is composed chiefly of *creasol* ($C_8H_{10}O_2$). In addition, creasote is said to contain *guaicol*, $C_7H_8O_2$, and phlorol, $C_8H_{10}O$. Inasmuch as true creasote is sometimes substituted by crude carbolic

acid, we give the following as distinctions which have heretofore been offered by various authorities, although true creasote differs very much when derived from various woods.

CREASOTE.	CARBOLIC ACID.
Not soluble in glycerin.	Soluble in glycerin.
Does not alter the appearance of collodium.	Precipitates gun cotton from collodium as a gelatinous mass.
Yields a green color with alcoholic solution of ferric chloride.	Yields a brown color with alcoholic solution of ferric chloride.

ETHERS.

The basylous hydrocarbon radicals do not (as we have said) exist singly in a free state, but they behave like metals, and form salts by combination with acid radicals. The salts formed by oxygenated acids are known as *compound ethers*; the salts of iodine, bromine, etc., as *haloid ethers*; the oxides of these unsaturated radicals are known as *simple ethers*, and the *hydroxides* are *alcohols*. Thus, the basylous radical *ethyl*, C_2O_5, which acts as a monad, by combination with—

Hydroxyl (OH) yields common alcohol, C_2H_5OH.
Iodine (I) yields iodide of ethyl, C_2H_5I.
Acetic acid radical ($C_2H_3O_2$) yields acetate of ethyl, $C_2H_5.C_2H_3O_2$.
Oxygen (O) yields oxide of ethyl, $(C_2H_5)_2O$.

Of the above—

Common alcohol is ethyl hydroxide, $C_2H_5.OH$.
Ethyl iodide is a haloid ether, $C_2H_5.I$.
Ethyl acetate is a compound ether, $C_2H_5.C_2H_3O_2$,
Ethyl oxide is a simple ether, $(C_2H_5)_2.O$.

The common sources of the unsaturated hydrocarbon radicals are the alcohols, and from the alcohols we produce the corresponding ethers. Of the many known ethers, a few only of the most common are employed in medicine.

Simple Ethers.—Of these, two are of interest to pharmacists and physicians. They are, as we have said, oxides of unsaturated basylous hydrocarbon radicals.

Ether.—Formula, $(C_2H_5)_2O$.

Synonyms.—Sulphuric ether. Ethyl oxide. Ethyl ethylate.

Ether is made by distilling a mixture of alcohol and sulphuric acid. The name, *sulphuric ether* was derived from this fact, but pure ether is free from any sulphur compound. In theory, the use of a limited supply of sulphuric acid should produce an unlimited amount of ether, as the acid is not consumed. In practice, however, it is found that the water produced by dehydration of the alcohol, together with the water and other impurities present in ordinary alcohol, interferes with and at last stops the operation. The reaction which takes place when ether is formed may be explained as follows. Alcohol and sulphuric acid first give rise to a molecule of water and a molecule of ethyl-hydrogen sulphate, as explained by the equation:—

$$\underset{\text{Alcohol.}}{C_2H_5OH} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Ethyl hydrogen sulphate.}}{C_2H_5HSO_4} + \underset{\text{Water.}}{H_2O}.$$

The molecule of ethyl-hydrogen sulphate then reacts with a fresh molecule of alcohol, the result being a molecule of ether (ethyl oxide) and a molecule of sulphuric acid, as follows:—

$$\underset{\text{Ethyl-hydrogen sulphate.}}{C_2H_5HSO_4} + \underset{\text{Alcohol.}}{C_2H_5OH} = \underset{\text{Ether.}}{(C_2H_5)_2O} + \underset{\text{Sulphuric acid.}}{H_2SO_4}.$$

Pure ether has the specific gravity of 0.712 or 0.710, boils at 93° or 94° F., and evaporates rapidly. It burns quickly and the vapor mixed with air or oxygen explodes when ignited. The Pharmacopœia recognizes, "*stronger ether*," of specific gravity not exceeding 0.728, and "*ether*," 0.750. These preparations contain alcohol. The last named is ordinary ether. It should be neutral to litmus, and evaporate without residue.

Methylic Ether.—Formula, $(CH_3)_2O$.

Synonyms.—Methyl oxide. Methoxyl-methane.

Mix, in a flask, by slowly stirring the acid into the alcohol, 1.3 parts of methylic alcohol with two parts of

sulphuric acid. Then connect the flask by means of a tube passing into a vial of alcohol. Upon heating the flask gently, a current of *methylic ether* will flow through the tube and into the alcohol where it is dissolved. In the foregoing reaction sulphuric acid abstracts the elements of water from the methylic alcohol as shown by the equation,

$$2CH_3OH - H_2O = (CH_3)_2O.$$

Methylic ether is a gas at ordinary temperatures, and from this reason we may not expect it to come into general use excepting in the form of solution. It is colorless, has an ethereal odor, and dissolves in water, alcohol or sulphuric acid.

Haloid Ethers.—These are combinations of unsaturated hydrocarbon radicals with members of the haloid group of elements. They may be considered as the result of displacement from saturated hydrocarbon radicals of one or more atoms of hydrogen by the haloid elements. Or, they may be classed as derivatives of alcohols by the displacement of hydroxyl (OH).

Hydrochloric Ether.—Formula, C_2H_5Cl.

Synonyms.—Ethyl chloride. Chlorethane.

This is best made by saturating ordinary alcohol with chlorine gas, and then distilling the mixture at a low temperature. The gas is conducted through a wash-bottle which contains a little warm water, and from thence to an ice-cold receiver. After condensation the ether is to be agitated with some fragments of fused calcium chloride, decanted and then redistilled. Hydrochloric ether is a thin volatile liquid, of specific gravity 0.921. It boils at 12.5° C. (54.5° F.) and is soluble in ten parts of *cold* water. This ether must not be confused with the so-called *chloric ether* which is simply a dilute solution of choroform in alcohol.

Hydrobromic Ether.—Formula, C_2H_5Br.

Synonyms.—Ethyl bromide. Bromethane.

This is made by distilling a mixture of amorphous phosphorus 1 part, bromine 8 parts and alcohol 32 parts. The

gas is purified by passage through a wash-bottle which contains a little warm water, and then is condensed in an ice-cold receiver. It is rather unsafe to attempt the preparation of this ether, unless the operator is experienced. Hydrobromic ether is volatile, heavier than water and has an ethereal penetrating odor and taste.

Hydriodic Ether.—Formula, C_2H_5I.

Synonyms.—Ethyl iodide. Iodethane.

This ether is made upon the principle of the production of hydrobromic ether, iodine being substituted for the bromine of that experiment. The reaction is explained upon the assumption that iodide of phosphorus is first produced, and this reacting with the alcohol produces an acid of phosphorus and ethyl iodide. Hydriodic ether is only made by such as have conveniences that guard against explosions and combustion. It is colorless, of specific gravity 1.9, has an ethereal odor, and decomposes upon exposure to light.

Ethene Chloride.—Formula, $C_2H_4Cl_2$.

Synonyms.—Ethylene chloride. Dutch liquid.

This ether is formed when olefiant gas (ethene) and chlorine gas are mixed. It has an oily appearance, and possesses a fragrant, ethereal odor.

Chloroform.—Formula, $CHCl_3$.

Synonyms. — Methenyl chloride. Formyl chloride. Chloroformum.

This, in medicine, is more extensively employed than any other haloid ether. It may be made by a variety of processes; for experimental purposes the formation of chloroform may be best illustrated by mixing a solution of chloral hydrate with a solution of caustic potash, formate of potassium, chloroform and water resulting as follows:—

$$\underset{\text{Chloral hydrate.}}{C_2HCl_3O.H_2O} + \underset{\text{Caustic potash.}}{KOH} = \underset{\text{Chloroform.}}{CHCl_3} + \underset{\text{Potassium formate.}}{CHKO_2} + \underset{\text{Water.}}{H_2O}.$$

The chloroform separates as a layer at the bottom of the vessel. Upon a manufacturing scale chloroform is made by distilling a mixture of chlorinated lime, water and alcohol. A complicated reaction ensues, several products arise, including chloroform, and it is unnecessary for us to consume space with the theoretical arguments that have been advanced upon this subject. Chloroform dissolves in pure alcohol, and in ether, in all proportions, and gives a sweetish taste to water. It is not readily inflammable, and after evaporation from glass or porcelain, should leave a clean surface. It should not alter the color of litmus paper, and when shaken with water should scarcely decrease in bulk (absence of alcohol). The specific gravity of pure chloroform is about 1.50 at 60° F. (15.5° C.) but as pure chloroform is unstable, the Pharmacopœia admits the admixture of alcohol, sufficient in amount to reduce its gravity to 1.480. Chloroform is colorless, and should neither impart color to, nor be colored by sulphuric acid when equal volumes are agitated and permitted to stand in contact for twenty-four hours. Chloroform has a sweet, aromatic taste, and an ethereal odor. It is an excellent solvent for some alkaloids. Chlorofom in overdoses is a poison. The liquid should be at once removed from the stomach by means of a stomach pump, the face of the patient exposed to a current of cold air, and artificial respiration conducted. Some recommend stimulants. In cases of poisoning by chloroform vapor, the above treatment is to be employed, *excepting* the stomach pump.

Iodoform.—Formula. CHI_3.

Synonym.—Methenyl iodide.

Iodoform is a haloid ether similar in composition to chloroform, iodine being substituted for chlorine. It is easily made by mixing one part ot iodine with two parts of alcohol, and with constant stirring, slowly adding a saturated solution of hydrate of potassium until the mixture is decolorized. The result is then to be poured into ight parts of distilled water, and the precipitate col-ed on a filter paper and dried by exposure to the air. form is a lemon-yellow substance of a persistent

iodine or saffron-like odor. It is soluble in alcohol, oils, and ether, and distills with the vapor of boiling water. Iodoform should completely dissolve in boiling alcohol, and should scarcely lose weight bydigestion with cold water. If it conforms to these tests it is not likely adulterated.

Compound Ethers. (Ethereal salts).

These constitute a large class, embracing the salts which are produced by the combination of organic basylous radicals with acid radicals. The formation of a compound ether from alcohol and an acid may be illustrated as follows: —

$$\underset{\text{Alcohol.}}{2C_2H_5OH} + \underset{\text{Oxalic acid.}}{H_2C_2O_4} = \underset{\text{Ethyl oxalate (oxalic ether)}}{(C_2H_5)_2C_2O_4} + \underset{\text{Water.}}{2H_2O}.$$

Practically, indirect processes are usually required to influence the economical combination of the basylous and the acidulous radicals.

Nitrous Ether.—Formula, $C_2H_5.NO_2$.

Synonym.—Ethyl nitrite.

If alcohol and nitric acid are gently warmed together, a complicated reaction ensues, aldehyd, nitrous ether and water resulting. If copper turnings and sulphuric acid are added, less aldehyd is said to result. In this case the nitric acid oxidizes the copper, which unites with the sulphuric acid to form sulphate of copper. The nitrite radical thus freed at once attacks the alcohol, nitrous ether resulting. The reaction is usually explained according to the equation which follows:—

$$\underset{\text{Alcohol.}}{C_2H_5OH} + \underset{\text{Nitric acid.}}{HNO_3} + \underset{\text{Sulphuric acid.}}{H_2SO_4} + \underset{\text{Copper.}}{Cu} =$$
$$\underset{\text{Nitrous ether.}}{C_2H_5NO_2} + \underset{\text{Copper sulphate.}}{CuSO_4} + \underset{\text{Water.}}{2H_2O}.$$

Nitrous ether is a thin, straw colored liquid with an ethereal odor. It is inflammable and its vapor if mixed with air explodes when lighted. This ether is not permanent, decomposing spontaneously and becoming quickly acid. It may be preserved by dilution with alcohol, and such is the officinal spirit of nitrous ether.

Spirit of Nitrous Ether, (Sweet spirit of nitre).

This was once made by distilling nitre (potassium nitrate), sulphuric acid and alcohol together, thus deriving the name "spirit of nitre." It is now made by the Pharmacopœial process by distilling a mixture of nitric acid four and one-half troy ounces, sulphuric acid three and one-half troy ounces, copper turnings two troy ounces, and alcohol twenty fluid ounces—the distillate afterward being mixed with 92 fluid ounces of alcohol. This spirit should contain five per cent. of the ether produced by the foregoing reaction. It should have the specific gravity 0.837, and it should not effervesce when to it a crystal of bicarbonate of potassium is added. The odor is ethereal, the color yellowish in large bulk, but it is colorless in small or ordinary amounts. The spirit of nitrous ether of the market is very unreliable, large amounts of it being mixed with water and sold under the names "4 F. and 3 F. spirit of nitre" the first term seems to signify *water*, the second, *more water*. If a pure solution of nitrous ether in alcohol be mixed with its bulk of chloroform or castor oil, the result will be transparent and colorless. If the watery mixture be tested in the same manner the result will at first be milky and then two layers will form.

Acetic Ether.—Formula, $C_2H_5.C_2H_3O_2$

Synonym.—Ethyl acetate.

This ether is easily made by distillation of a mixture of acetate of sodium, alcohol, and sulphuric acid. During the reaction water and acetic ether are formed and distill, while acid sulphate of sodium remains in the retort. The easiest process for making acetic ether, is that of passing carbon dioxide into an alcoholic solution of potassium acetate. The result is distilled. After both processes the ether should be purified by digestion with chloride of calcium, and carbonate of calcium, then decanted and redistilled. Acetic ether has an agreeable odor, mixes with alcohol and ether in all proportions, but dissolves only to the extent of about ten per cent. in water. It is not used as medicine, but is an excellent solvent of cantharidine.

Often it will produce instantaneous solution of gun cotton, if added in small amount to the collodium mixture, when without it the cotton refuses to dissolve, forming a gelatinous mass.

Amyl Nitrite.—Formula, $C_5H_{11}.NO_2$.

Synonym.—Nitrite of amyl.

This is made by mixing amylic alcohol (purified) with an equal bulk of nitric acid, and carefully distilling at a temperature not exceedeng 100° C. (212° F.) The distillate is then agitated with an aqueous solution of potassium carbonate, and the overlying ethereal solution is decanted and redistilled; that portion which distills between the temperatures of 95° C. and 100° C. being reserved. Nitrite of amyl is of the specific gravity 0.877, has a yellowish color, and a peculiar ethereal odor. It is insoluble in water but mixes with alcohol in all proportions.

Glycerides.—These are combinations of glyceryl (C_3H_5) with the radicals of certain fatty acids, such as oleic acid, stearic acid, etc. These combinations give rise to the various fats and oils, and thus we call oils and fats compound ethers. Some of these ethers are solid at ordinary temperatures, others are liquid. Natural fats consist of mixtures of the two classes in various quantities, and in proportion as the hard or soft ethers preponderate, hard or soft fats result.

Olein.—Formula, $C_3H_53C_{18}H_{33}O_2$.

Synonym.—Glyceryl trioleate.

This body constitutes the bulk of most liquid fats or oils. It is a combination of the acidulous radical ($C_{18}H_{33}O_2$) of oleic acid with glyceryl (C_3H_5). The production of this substance from propenyl alcohol (glycerin) and oleic acid may be illustrated as follows:—

$$\underset{\text{Glycerin.}}{C_3H_53OH} + 3\underset{\text{Oleic acid.}}{HC_{18}H_{33}O_2} = \underset{\text{Olein.}}{C_3H_5(C_{18}H_{33}O_2)_3} + \underset{\text{Water.}}{3H_2O}.$$

When olein is heated in contact with water under a steam pressure, decomposition results, and both oleic acid and glycerin result. Oleic acid combines with metallic oxides to form salts which are known as *oleates*. Olein is

of great use in the arts, and is used in medicine for the preparation of glycerin, plasters, cerates, and ointments. The officinal *liniment of ammonia*, a mixture of ammonia water one part and olive oil two parts, is an imperfect *oleate of ammonium*. Drying oils, such as linseed oil, contain an oil (linolein) of the composition $C_3H_5 3C_{18}H_{27}O_2$. Castor oil contains *recinolein*, $C_3H_5 3C_{18}H_{33}O_3$.

Stearine.—Formula, $C_3H_5 3C_{16}H_{31}O_2$.

Synonym.—Glyceryl tristearate.

This is a hard fat, constituting the bulk of mutton tallow, cocoa butter, etc. When pure it is white, and presents a crystalline appearance. Stearine is employed in a greater or less degree of purity for making candles and animal soaps.

Palmitin.—Formula, $C_3H_5 3C_{16}H_{31}O_2$.

Synonym.—Glyceryl tripalmitate.

This, in small quantity, is found in most solid fats, but in large proportion in palm oil, of which it constitutes the solid portion. The so-called *margaric acid* is a mixture of palmatic acid ($HC_{16}H_{31}O_2$.) and other fatty acids.

ALDEHYDS.

These bodies are the result of oxidation of alcohol. We have shown heretofore how ordinary aldehyd may be derived from ordinary alcohol, and upon the same principle, other aldehyds may be obtained by the action of various oxidizing agents upon the alcohols. The name aldehyd is applied to the aldehyd of common alcohol and it is also known as acetic aldehyd. This is not employed in medicine, unless it be as a contamination of spirit of nitrous ether. Aldehyds form precipitates with hydrogen potassium sulphite (acid sulphite of potassium).

Chloral.—Formula, C_2HCl_3O.

Synonym.—Trichloraldehyd.

This body is made from anhydrous alcohol by the action of dry chlorine gas, whereby the alcohol first changes

to aldehyd; this by displacement of three atoms of hydrogen forms chloral. Thus:—

$$\underset{\text{Aldehyd.}}{C_2H_3O.H} + \underset{\text{Chlorine.}}{2Cl_3} = \underset{\text{Chloral.}}{C_2HCl_3O} + \underset{\text{Hydrochloric acid.}}{3HCl.}$$

The above is the simplest theoretical explanation of the changes which result from the passage of chlorine into pure alcohol. In practice several incidental products arise, and actually the reaction is more complex. Chloral is not made to any extent in this country. Its preparation is tedious and expensive. Our supply is obtained from other countries, where chlorine is a waste product. Chloral is an oily liquid, colorless, evolving vapors which are very irritating to the eyes. It is employed for the purpose of making—

Chloral Hydrate.—$C_2HCl_3O.H_2O$.

Synonyms.—Chloral. Hydrate of chloral.

When the oily liquid chloral, is mixed with a small amount of water, the mixture solidifies as a mass of crystals of dazzling whiteness. The odor is pleasant and peculiar. The chlorine in chloral hydrate refuses to precipitate nitrate of silver solution, thus differing from metallic chlorides. Chloral hydrate sublimes without decomposition, and if rubbed with camphor gum produces a liquid. Water dissolves about three and one-half times its weight of chloral hydrate, and if the solution be diluted, acidulated with sulphuric acid, and colored slightly with solution of potassium permanganate, it will not decolorize in several hours, if the chloral hydrate was pure. When solution of choral hydrate is mixed with solution of a caustic alkali, chloroform and an alkaline formate result. Theoretically the yield of chloroform should be 72.2 per cent. of the weight of the chloral hydrate employed, but practically the yield is about 70 per cent. Chloral hydrate will not turn cold sulphuric acid brown, and will not evolve red fumes with nitric acid if they be gently warmed together. This distinguishes the true chloral hydrate from *chloral alcoholate*, for the latter article turns brown with sulphuric acid, and evolves red fumes

when heated with nitric acid. The blood of animals is of alkaline reaction, and from this fact may be explained the slow liberation of chloroform after the administration of chloral hydrate, together with its continued and lasting effect as compared with chloroform. Strychnine has been recommended as an antidote for poisoning by this drug, but it must be used with caution. At once empty the stomach with a stomach pump, then freely administer strong coffee or tea. Use the antidotes recommended for poisoning by chloroform and by morphine.

Benzoic Aldehyd. (C_7H_5OH)—

Oil of bitter almonds, and *cinnamic aldehyd* (C_9H_7OH), oil of cinnamon, are described under the essential or volatile oils. The aldehyds we have named are all that are used in medicine.

Organic Acids.

These are considered as the results of the oxidation of alcohols. In the first place the alcohol molecule parts with a molecule of hydrogen, water and an *aldehyd* resulting. The aldehyd, upon exposure, absorbs an atom of oxygen and forms the acid. Thus we illustrate the formation of acetic acid from ordinary alcohol—

$$\underset{\text{Alcohol.}}{2C_2H_5HO} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Aldehyd.}}{2C_2H_4O} + \underset{\text{Water.}}{2H_2O}.$$

The above is the first step, and—

$$\underset{\text{Aldehyd.}}{2C_2H_4O} + \underset{\text{Oxygen,}}{O_2} = \underset{\text{Acetic acid.}}{2C_2H_4O_2},$$

is the second part of the reaction. The organic acids employed in medicine were discovered and named before their connection with the alcohols was determined, or even before the several alcohols were known. Consequently the names of the acids and the connected alcohols do not correspond. We have included the organic acids, with our list of acids (salts of hydrogen) as we believe the practical study of each compound is at present of value to active pharmacists, physicians and medical students, rather than the advanced views of theoretical chemists.

CARBO-HYDRATES.

These bodies constitute the bulk of most plants. They are named *carbo-hydrates* because each contains 6 or 12 atoms of carbon united with the exact proportion of oxygen and hydrogen which is requisite to form a definite number of molecules of water, thus:—

Cellulose.....$C_6H_{10}O_5$.
Cane sugar.....$C_{12}H_{22}O_{11}$.
Grape sugar.....$C_6H_{12}O_6$.

The carbo-hydrates are at present divided into three isomeric classes, those of interest in medicine being as follows:—

AMYLOSES.		SACCHAROSES.	
Cellulose	$C_6H_{10}O_5$.	Cane Sugar	$C_{12}H_{22}O_{11}$.
Starch	$C_6H_{10}O_5$.	Milk Sugar	$C_{12}H_{22}O_{11}$.
Dextrin	$C_6H_{10}O_5$.		

GLUCOSES.			
Grape Sugar	$C_6H_{12}O_6$.	Fruit Sugar	$C_6H_{12}O_6$.

Amyloses.—This class of carbo-hydrates includes all which are of the general formula $C_6H_{10}O_5$. They are isomeric or polymeric† with each other, but their molecular weights have not been accurately determined. The name of this class of compounds is derived from *amylum* (starch), the principal constituent of the group.

Starch.—Formula, $C_6H_{10}O_5$ (or $C_{18}H_{30}O_{15}$).

Synonym.—Amylum.

Starch is a white granular substance, and is found throughout the entire vegetable kingdom, but most abundantly in certain tubers and in seeds. Starch is insoluble in cold water, but in boiling water the granules

† Polymeric molecules differ in molecular weight, but they contain similar atoms united in the same proportion. Thus the difference between the molecule of acetic acid and the molecule of glucose, is in the *number* of atoms, and not in *percentage*. Glucose $C_6H_{12}O_6$, contains three times the number of atoms, and the same elements contained in a molecule of acetic acid, $C_2H_4O_2$.

swell and burst, producing a thick mucilaginous liquid, known as starch paste. If starch be boiled with a mixture of sulphuric acid and water, it is converted into glucose and dextrin. If desirable, glucose can be alone produced. If the operator be skillful, only a few moments are required to change every trace of starch into glucose, *free* from dextrin, although most authorities state that prolonged boiling is necessary. If starch be acted upon by nitric acid, oxalic acid is formed. If a few drops of starch paste be mixed with a little water, and tincture of iodine be added, a deep blue color results. This reaction is a most delicate test for either starch or iodine. It must be remembered in this case that the starch granules are to be burst by hot water, thus forming starch paste.

Dextrin.—Formula, $C_6H_{10}O_5$.

Dextrin may be made, as we have said, by boiling starch with diluted sulphuric acid, glucose also resulting, thus:—

$$\underset{\text{Starch.}}{(C_6H_{10}O_5)_3} + \underset{\text{Water.}}{H_2O} = \underset{\text{Glucose}}{C_6H_{12}O_6} + 2\underset{\text{Dextrin.}}{C_6H_{10}O_5}.$$

In practice dextrin is made by heating dry starch to 350° F.

Dextrin is a yellowish powder, possessing a peculiar odor which is very disagreeable to some persons. It dissolves in water, forming a sticky solution much employed as a mucilage. The coating applied to the back of postage stamps seems to consist of nearly pure dextrin.

Cellulose.—Formula, $C_6H_{10}O_5$.

Synonyms.—Lignin. Cellulin.

Cellulose is the substance which forms the cell walls of plants. Familiar examples of cellulose are cotton, and the fiber of flax (linen). Water does not dissolve cellulose, but a solution of basic carbonate of copper in ammonia water dissolves it readily, and from this solution an acid precipitates it as white flakes. If cellulose be digested with cold sulphuric acid a gummy substance resembling dextrin results, and this produces glucose if it be boiled with dilute sulphuric acid. Cellulose is of great value in the arts as a fabric producer, and cotton is used in pharmacy to make—

Nitro-cellulose, or *gun cotton.* This is officinal under the name *pyroxylin,* and is made by treating cotton with strong nitric acid, whereby either one, two, or three atoms of hydrogen are replaced by nitril (NO_2), thus:—

Mononitro-cellulose.... .....$C_6H_9(NO_2)O_5$.
Dinitro-cellulose.....$C_6H_8(NO_2)_2O_5$.
Trinitro-cellulose .....$C_6H_7(NO_2)_2O_5$.

The last of the series is true gun cotton, being the most explosive. The second is preferred for making collodium as it dissolves freely in ether. Officinal collodium is made by mixing stronger ether, twelve fluid ounces, with stronger alcohol, three fluid ounces, and dissolving in the mixture two hundred grains of pyroxylin. In view of the fact that pyroxylin is consumed in the photograph business in large amounts, under the name *negative cotton,* and is supplied to the trade at low prices, it is not economic to prepare the cotton on a small scale.

Gums.

The class of organic bodies properly known as gums, agree somewhat in composition with starch, and may consistently be considered in this place. In commerce a number of bodies, such as *catechu, kino,* etc., are called gums, but they are the dried juices of plants, or dried solid extracts. We are compelled to notice such as are of interest in medicine, and shall include them with gums.

Gum Arabic.—This article, known also as gum acacia, gum Arabin, or as acacia, exudes in tears from certain species of acacias in Egypt and Arabia. It contains mainly a soluble principle, *arabin* ($CaC_{12}H_{20}O_{11}$), which is in reality a calcium salt of gummic (Arabic) acid, $H_2C_{12}H_{20}O_{11}$. It is customary, therefore, for some to speak of gum Arabic as *gummate* of calcium. If hydrochloric acid be added to a solution of gum Arabic, chloride of calcium and gummate of hydrogen (gummic acid) result, thus:—

$$CaC_{12}H_{20}O_{11}+2HCl=CaCl_2+H_2C_{12}H_{20}O_{11}.$$

Gummic acid is isomeric with cane sugar, and when it is boiled with dilute sulphuric acid, gives rise to glucose. If

a few drops of ferric sulphate solution be added to a solution of gum Arabic, a gelatinous mass results.

Gum Tragacanth.—This gum is the exudation of the trunk of the Astragalus verus. It is white, does not dissolve in cold water, but swells when covered with water, and forms a translucent gelatinous mass. It is used extensively as a paste, and this is prepared by pouring about sixteen parts of cold water upon one part of gum tragacanth, and permitting the mixture to stand twenty-four hours. The chief constituent of tragacanth is *bassorin.* Tragacanth is not liable to adulteration, but very inferior brown qualities are upon the market.

Catechu.—This substance, often known as *gum catechu,* is a solid extract of the wood of acacia catechu. It occurs in dark brown masses, more or less impure from mixtures of twigs, chips, and gravel. Catechu is astringent, yields when treated with cold water mostly *catechutannic* acid. The insoluble matter contains *catechin* (catechuic acid), and it may be extracted by boiling water, and purified by repeated crystallizations. Tincture of catechu gelatinizes, as is the case with tincture of kino, and other tinctures of the vegetable kingdom which contain certain red tannates.

Kino is the dried juice of an India tree. It is an astringent extract, but not a true gum. In some respects it resembles catechu, but it has no catechuic acid, and contains a pectin compound. Tincture of kino is used extensively in medicine.

Aloes.—The juices of several species of aloe, furnish, when dried, more or less brittle and dark colored extractive masses, known as "aloes." Aloes is a mixture of substances, and its value is not generally estimated by chemical constituents.† It should dissolve in diluted alcohol.

† There are several varieties of aloes in commerce, distinguished as Barbadoes aloes, Socotrine aloes, Zanzibar aloes, Cape aloes, Natal aloes, etc. These contain substances called aloins, and each variety differs chemically, and in purging properties.

Saccharoses.—This section of the carbo-hydrates, embraces cane-sugar and other bodies of the general formula $C_{12}H_{22}O_{11}$. They are regarded by some as (theoretical) poly glucosic alcohols, in which two or more molecules of glucose unite. Such combinations are accompanied by the elimination of water molecules, one less than the number of glucose molecules. The molecular weights of the saccharoses have never been established.

Sugar.—Formula. $C_{12}H_{22}O_{11}$.

Synonyms.—Cane sugar. Sucrose. Saccharose. Beet root sugar. Saccharum album.

Cane sugar is found in all parts of the vegetable kingdom, but sugar-cane, beet-root, and the sugar-maple tree supply the markets of the world. The processes employed in making cane sugar, are familiar to most persons, and are explained at length by works upon technical chemistry. The purest form of cane sugar is that known as loaf or crushed sugar. Cane sugar melts if heated forming "barley sugar," and at a high heat decomposes, turns black, and the result is known as *caramel*, or burnt-sugar, a preparation employed in coloring liquors. Some metallic salts decompose when boiled with sugar solution, the metals precipitating. By fermentation of sugar alcohol first results, then acetic acid. Warm nitric acid forms saccharic acid ($C_6H_{10}O_8$) with sugar; boiling nitric acid produces oxalic acid. Sulphuric acid chars sugar, abstracting the elements of water. Diluted sulphuric acid, if boiled with sugar, produces *inverted sugar*. "Fehling's solution" is not affected when boiled with a fresh solution of cane sugar, which fact serves to distinguish cane sugar from the most common adulterant, glucose.

Milk Sugar.—Formula, $C_{12}H_{22}O_{11}.H_2O$.

Synonyms.—Saccharum lactis. Lactin.

Milk sugar is extensively produced in some parts of Europe, but it is not made in this country, notwithstanding the bountiful supply of milk in certain sections. Milk sugar is made by simply evaporating whey to crystallization, and afterward purifying the crystalline masses by

[illegible] them in water, filtering the solution through [illegible], and then recrystallizing. Sugar of milk [illegible] is slightly sweetish, odorless, and grits between [illegible]. It is insoluble in ether, slightly soluble in [illegible], soluble in from five to six parts of cold water, [illegible] about three parts of boiling water. It reduces the [illegible] of "Fehling's solution," and is estimated in the [illegible] manner. When a solution of milk sugar is mixed with powdered chalk and decayed cheese, and digested [illegible] weeks in a warm situation, *lactic acid* results. [illegible] boiling milk sugar with dilute sulphuric acid, a [illegible]ance (galactose) isomeric with glucose results. Nitric [illegible] oxidizes milk sugar, mucic acid being formed; an [illegible] (saccharic acid) is produced under like conditions by most other sugars. Milk sugar has no medicinal value, but as an inert substance is used in the preparation of [illegible] where the object is the division of active [illegible]. *Maltose* is the sugar of malt resembling glucose in many respects. It is said to be an isomer of milk sugar. It is not used in medicine.

Glucoses.

These are of the general formula $C_6H_{12}O_6$, and of the members composing this group which are known at the present day, two or three only are of interest in pharmacy, or in medicine. The glucoses give a precipitate of cuprous oxide when boiled with Fehling's solution, and their solutions in cold water are not charred by sulphuric acid.

Glucose.—Formula, $C_6H_{12}O_6$.

Synonyms.—Dextrose, Grape sugar, Dextro-glucose, Corn-sugar.

This is the important member of the group; it is found in most fruits and plants, and constitutes a large percentage of honey. This sugar also appears more or less [illegible] in the urine of diabetic persons, and its detection [illegible] conditions is a necessity which is recognized [illegible] chemists of the present day. Glucose is now [illegible] on a dense amounts in the Western cities of this [illegible] from the starch of corn, and under the name

glucose, or *corn sugar*, is employed in preserving works, confectioneries, candy-making, etc. Its sweetening powers are, however, very inferior to those of cane sugar. Any of the natural glucosides yield glucose when boiled with dilute sulphuric acid, and all substances containing cellulose, by digesting with strong sulphuric acid, and subsequent boiling with diluted sulphuric acid, give glucose in proportion to the cellulose present. Owing to the fact that the manufacture of glucose has now become so important a branch of industry, and that its mode of preparation is explained so fully in various works upon technical chemistry, we do not feel called upon to consume space with the practical part of its preparation. Theoretically the starch of the corn is first converted into a mixture of dextrine and glucose, a molecule of water being taken up by three molecules of starch, usually expressed as follows:

$$\underset{\text{Starch.}}{3C_6H_{10}O_5} + \underset{\text{Water.}}{H_2O} = \underset{\text{Dextrine.}}{2C_6H_{10}O_5} + \underset{\text{Glucose.}}{C_6H_{12}O_6}.$$

The dextrine by further combination with water becomes glucose—

$$\underset{\text{Dextrine.}}{2C_6H_{10}O_5} + \underset{\text{Water.}}{2H_2O} = \underset{\text{Glucose.}}{2C_6H_{12}O_6}.$$

In practice starch may be apparently converted *directly* into glucose. According to the author's experiments, if the conversion is not direct, it is so nearly so, that there is no intermediate production of dextrine, which necessitates prolonged boiling. Indeed the object is to *guard against* the formation of dextrine, a point easily accomplished, and large amounts of starch may be converted in a short period entirely into glucose. Glucose dissolves in 140 parts of water, crystallizes in small, warty aggregations or cubes, and these dissolve in concentrated sulphuric acid without blackening (cane sugar turns black). For the detection of glucose in the urine, see the section of this work which is devoted to examination of urine.

Fruit Sugar.—Formula, $C_6H_{12}O_6$.

Synonyms.—Honey sugar. Inverted sugar.

When cane sugar is warmed with dilute sulphuric acid it is transformed into equal parts of glucose and levulose, thus:—

$$\underset{\text{Cane sugar.}}{C_{12}H_{22}O_{11}} + \underset{\text{Water.}}{H_2O} = \underset{\text{Levulose.}}{C_6H_{12}O_6} + \underset{\text{Glucose}}{C_6H_{12}O_6}.$$

The cane sugar is said to be *inverted*, and the mixture is known as *fruit sugar* or *honey sugar*, being found in many fruits and in honey. Levulose, one of the products, is as sweet as cane sugar, uncrystallizable, and it turns a ray of polarized light to the left.

Amines.

This class of bodies includes substances which are built upon the type of ammonia, by substitution of basylous radicals for the hydrogen of ammonia radicals. If the hydrogen of one molecule of ammonia be displaced, we have a *monamine*. If the amine is derived from two molecules of ammonia, it is called a *diamine*. If from three molecules of ammonia it is a *triamine*, etc. It is unnecessary for us to enter into the theoretical study of this class of bodies, and interested parties are referred to the late works which are devoted to a systematic review of theoretical investigations.

The amines are strong bases and unite with acids like ammonia, forming salts. The organic bases known as alkaloids are classed with bodies of this description, and are called *compound ammonias*. As a class, however, the molecular construction of the alkaloids is unknown.

Trimethylamine.—Formula, $N(CH_3)_3$.

Synonym.—Propylamine.

This body has long been sold in commerce under the name propylamine. True propylamine (C_3H_9N) is an isomer of trimethylamine, and is made by hydrogenation of cyanide of ethyl (C_3H_5N). Trimethylamine is made by distilling fish brine with an alkali, and the distillate which contains crude trimethylamine is afterward purified. The process is tedious and uncalled for in a work of this description. *Trimethylamine* is a liquid, colorless, of a strong herring-like odor, strongly alkaline, forming with acids, salts which are freely soluble in both alcohol and water. Trimethylamine is soluble in water, ether and in alcohol; it is very volatile.

Chloride of propylamine of the market is in reality the chloride of trimethylamine, a soluble and deliquescent salt formed by direct combination of hydrochloric acid and trimethylamine.

ALKALOIDS.

The term "alkaloid" is applied to organic bases of the vegetable and animal kingdoms. Sometimes alkaloids are called *organic alkalies*. There has been much confusion regarding the nomenclature of the alkaloids. In this work preference is given to the termination *ine*. Thus, instead of berberin, or berberina, we say *berberine*, placing the other names among the synonyms. At present we are led to believe alkaloids are all derivatives of ammonia molecules, therefore we class them with the amines. When an alkaloid is heated with hydroxide of potassium, ammonia is evolved. Alkaloids are alkaline, and combine with acids to form salts. Some are strong bases, for example, berberine, others are very feeble, as caffeine, which refuses to unite with organic acids, excepting under certain conditions, and the salts of which, even with mineral acids, are decomposed by water. Alkaloids unite with acids wholly, and the act of combination does not give rise to water; thus they differ from the inorganic bases. For this reason the term *muriate*, or *hydrochlorate*, is applied to the combination of an alkaloid and hydrochloric acid, instead of the term *chloride*. Some alkaloids are volatile, others are fixed, and decompose upon heating. The volatile alkaloids are free from oxygen, while the fixed alkaloids all contain oxygen. A large number of alkaloids are known, and, as a rule, the alkaloids are active medicinal agents. Some alkaloids are the most powerful of poisons. As a rule, aqueous solutions of the alkaloids, or their salts, give precipitates with phosphomolybdate of ammonium, chloride of platinum (chloroplatinic acid, H_2PtCl_6) or solution of iodine in iodide of potassium solution. We shall only notice such alkaloids as are of interest from a therapeutical point.

Aconitine.—Formula (old), $C_{30}H_{47}NO_7$.

Synonym.—Aconitia.

This alkaloid is obtained from the root and leaves of *Aconitum napellus.* As found in market it is an amorphous, white powder, odorless, but very irritating to the mucous membranes of the nose, even if in very small amount. It is intensely acrid and burning to the taste, producing a sense of numbness; more soluble in cold than in hot water, and is soluble in alcohol and ether. The foregoing description is generally accepted, but it appears that the article named is a mixture of alkaloids. True *aconitine* should be crystalline, of the composition $C_{33}H_{43}NO_{12}$. From the fact, however, that some batches of root refuse to yield the crystalline alkaloid, and yield the amorphous freely, and that the aconitine of commerce is the amorphous, we have given that article preference. Aconitine is an active poison, and can not well be administered internally, $\frac{1}{50}$ of a grain producing poisonous effects.

Atropine.—Formula, $C_{17}H_{23}NO_3$.

Synonym.—Atropia.

This alkaloid exists in belladonna as a malate. It is white, soluble in water, alcohol, ether, glycerin, and chloroform, and by spontaneous evaporation of its solutions assumes the form of needle-like crystals. It dissolves colorless, in cold sulphuric acid, and the solution turns brown upon heating. The solution of atropine in cold nitric acid is yellowish, changing to orange by heating. Atropine is poisonous. It dilates the pupil of the eye, thus being opposite in effects to the alkaloid of calabar bean (physostigmine). The *sulphate of atropine* is very soluble in water, and this form of atropine is mainly used in medicine.

Beberine.—Formula, $C_{19}H_{21}NO_3$.

Synonyms.—Beeberin. Bebirin. Beberia.

This alkaloid is found in the bark of *Nectandra Rodiæi.* The pure alkaloid is white, very bitter, freely soluble in alcohol and in ether, but almost insoluble in cold water. It forms salts with acids. *Sulphate of beberine* is the ar-

ticle employed in medicine, the pure alkaloid being seldom used. Sulphate of beberine is not crystallizable and is soluble in water and alcohol. It appears in market as dark-brown scales, having been made by spreading a syrupy solution upon glass and scraping the glass when the solution has dried. Beberine must not be confused with *berberine* (hydrastine) the yellow alkaloid found so abundantly in the rhizome of hydrastis canadensis.

Berberine.—Formula (Perrins), $C_{20}H_{17}NO_4$.

Synonyms.—Berberina. Berberin. Hydrastin. Jamaicin. Xanthropicrit.

Berberine is employed most extensively of any alkaloid obtained from North American plants. It received the name from *Berberis vulgaris.* The market at present is supplied with the alkaloid as produced from the rhizome of Hydrastis Canadensis, and the name by which the alkaloid is known among physicians is *hydrastine.* Berberine is practically produced by the decomposition of sulphate of berberine through the agency of an alkali. As thus produced from rhizome of Hydrastis Canadensis, it is lemon-yellow, very soluble in water, less in alcohol, and is insoluble in ether. It is very bitter, crystallizes from solution in tufts of acicular needles, and combines directly with acids, forming salts which, as a rule, are less soluble than the alkaloid. Picrate of berberine is insoluble.

Muriate of Berberine.

This salt of the alkaloid is easily made by adding hydrochloric acid to tincture of Hydrastis Canadensis. It separates as yellow needles, and may be purified by washing with water, and then recrystallizing from solution in boiling water. Muriate of berberine is of a lemon yellow color, soluble in about 500 parts of cold water, almost insoluble in alcohol, insoluble in ether and in chloroform. Picric acid precipitates it from solution as picrate of berberine, but the hydrochloric acid refuses to be displaced by even the strong mineral acids. Oxide or phosphate of silver, if boiled with a solution of muriate of berberine, liberates the alkaloid.

Sulphate of Berberine.

This preparation is now the most common salt of the alkaloid. It is made by adding an excess of sulphuric acid to an alcoholic tincture of Hydrastis Canadensis, and then purifying the needle-like crystals in the manner we have directed for muriate of berberine. Sulphate of berberine is in the form of orange-yellow crystalline tufts, quite soluble in cold water, very soluble in boiling water, slightly in alcohol, and insoluble in ether and chloroform. When an alkali is added to the aqueous solution of sulphate of berberine, the salt is decomposed and berberine is liberated. The other salts of berberine may be easily made by adding the desired acid to solution of the alkaloid.

Brucine.—Formula, $C_{23}H_{26}N_2O_4$.

Synonyms.—Brucia. Brucin.

This alkaloid is obtained as a by-product after strychnine is crystallized from an extract of nux vomica. Brucine is white, very bitter, soluble in 800 parts of water, and insoluble in ether. It dissolves in alcohol and chloroform. Brucine forms a red solution with nitric acid, which turns yellow upon heating. This solution changes to purple-violet with solution of stannous chloride (solution of muriate of tin). Brucine is poisonous, but is less violent in action than strychnine.

Caffeine.—Formula, $C_8H_{10}N_4O_2.H_2O$.

Synonyms.—Caffin. Caffeidina. Caffeia. Theine. Theia.

This alkaloid is found in more or less abundance in tea, coffee, and guarana. It may be obtained simply by treating the powdered drug with boiling chloroform, filtering, and evaporating the filtrate to dryness. The mass is extracted with hot water, which dissolves the caffeine, and evaporation of the water leaves it pure. Caffeine is white, odorless, slightly bitter, and crystallizes in silky needles. It dissolves in about ten parts of boiling water and in sixty parts of cold water. Chloroform dissolves one-ninth its weight of caffeine. It is slightly soluble in ether, and soluble in twenty-one parts of alcohol. The solutions of

caffeine are neutral, the solution in water is precipitated by tannic acid. When the gas which is evolved from a mixture of hydrochloric acid and potassium chlorate, is passed into solution of caffeine in water, and the solution afterward evaporated, a blood-red residue results. This reaction is said to detect one part of caffeine in one thousand parts of water. Caffeine sublimes unchanged if heated. It forms salts with the strong mineral acids, but water decomposes them. Organic acids do not freely unite with caffeine. Citrate of caffeine is the salt most in demand. It is very unstable, decomposing under the influence of water or alcohol.

Alkaloids of the Cinchonas.

The many varieties and species of cinchona contain a number of alkaloids which resemble each other to an extent, and have more or less value as antiperiodics. These alkaloids are naturally combined with kinic, kinovic, and cinchotannic acids; and the barks of the cinchonas owe their virtues to such combinations. Some qualities of bark yield a large per cent. of alkaloids, others are almost devoid of these principles. The U. S. P. recognizes a bark containing two per cent. of alkaloids which are capable of forming crystalline salts. We can only notice the alkaloids which are in common use as medicines, and their salts.

Quinine.—Formula, $C_{20}H_{24}N_2O_2.3H_2O$.

Synonym—Quinia.

This is the most valuable of the cinchona alkaloids. The bark known as *Calisaya* yields it most abundantly. It is best produced from sulphate of quinine, and as follows:—Dissolve sulphate of quinine in fifty parts of cold, distilled water by the careful addition of a sufficient amount of sulphuric acid. Into this solution carefully stir ammonia water until it is in slight excess. This will form sulphate of ammonium, the alkaloid quinine being liberated in the form of a white precipitate. Filter the solution and dry the precipitated quinine by exposure to the atmosphere in a cold location. If the temperature be

"Agitate two grammes (30 grains) of sulphate of quinine with twenty cubic centimeters (5 fluid drachms) of distilled water at the temperature of 15° C. (59° F.) and filter after a maceration of half an hour. Add, carefully, to five cubic centimeters (1¼ fluid drachms) of this solution, in a test-tube, seven cubic centimeters (1⅔ fluid drachms) of ammonia water of specific gravity 0.96. Now close the mouth of the test-tube and invert the tube, there should be formed, either immediately or after a short time, a clear liquid."

In case the sulphate of quinine is contaminated with salts of the other cinchona alkaloids, the alkaloids remain insoluble, and the resulting fluid is not transparent.

Quinine Hydrochlorate. — Formula, $C_{20}H_{24}N_2O_2.HCl.3H_2O$.

Synonyms.—Muriate of quinine. Muriate or hydrochlorate of quinia.

This salt is seldom employed in medicine. It may be made by dissolving the alkaloid quinine in dilute hydrochloric acid, and by proper management may be obtained in silky crystalline tufts.

Quinine Tannate is made by mixing solution of bisulphate of quinine with solution of tannin. It precipitates as a buff colored magma, is almost insoluble in water, and scarcely possesses a bitter taste. The freshly precipitated article will dissolve to a considerable extent in glycerin, and such solution is almost tasteless. Other salts of quinine may be made by dissolving the alkaloid in the respective acids, but they are not of sufficient importance for individual notice in our work.

Cinchonidine.—Formula, $C_{20}H_{24}N_2O$.

Synonym.—Cinchonidia.

This alkaloid is isomeric with cinchonine, and like cinchonine is obtained from the mother liquor after crystallization of sulphate of quinine. It is sold in the form of sulphate $(C_{20}H_{24}N_2O)_2\ H_2SO_4.4H_2O$. It forms tufts of needle-like crystals which resemble sulphate of quinine. Sulphate of cinchonidine is soluble in alcohol, but insolu-

ble in ether. It does not form a green precipitate with chlorine water and ammonia, (distinction from quinine). The solution of sulphate of cinchonidine in water affords precipitates with solution of Rochelle salt, which distinguishes it from cinchonine (cinchonia). The sulphate of cinchonidine, until recent years almost unknown to the commercial world, is now sold in quantities as a substitute for the expensive salts of quinine.

Cinchonine.—Formula, $C_{20}H_{24}N_2O$.

Synonym.—Cinchonia.

This is the cheapest alkaloid of cinchona bark and recently has been in considerable demand as a tasteless anti-periodic. It is obtained from the mother liquor of the quinine manufacture. Cinchonine is insoluble in water and in ether. Muriate of cinchonine has been substituted for sulphate of quinine. This may be detected by dissolving the suspected salt in a little water acidulated with sulphuric acid, then adding ammonia water in slight excess, and washing and drying the precipitate. This should perfectly dissolve in ether if it be quinine. If it refuses to dissolve in ether, and after solution in very dilute sulphuric acid, refuses to precipitate upon the addition of Rochelle salt, it is cinchonine. If Rochelle salt produces a precipitate, *cinchonidine* is present. Solutions of cinchonine salts are very bitter and are not fluorescent.

Quinidine.—Formula, $C_{20}H_{24}N_2O_2$.

Synonym.—Quinidia.

This alkaloid is in little demand. It is white, isomeric with quinine, and, like cinchonine and cinchonidine, is obtained from the mother liquor after crystallization of quinine salts. Quinidine is sold as a sulphate, which dissolves readily in dilute sulphuric acid.

Chinoidine is a brownish mass left after all the crystallizable alkaloids of cinchona are separated from the solution. It consists of amorphous alkaloids, and is employed to a considerable extent in the form of pills. Chinoidine is tasteless, insoluble in water, but soluble in alcohol and dilute acids, and such solutions are bitter. The

other alkaloids of cinchona are of interest to experimental chemists only, as they are not used in medicine, excepting in combination.

Codeine.—Formula, $C_{18}H_{21}NO_3$.

Synonyms.—Codeia. Codein.

This alkaloid exists in opium, and is obtained from the mother liquor of the morphine manufacture. As it is seldom used in medicine, and is never prepared on a small scale, the mode of preparation is of no particular interest to the general student. Codeine is a white alkaloid, soluble in seventeen parts of hot and in eighty parts of cold water, soluble in ether and in alcohol. Codeine has little taste, but its salts are bitter. Solution of codeine (aqueous) is alkaline. Codeine is insoluble in solution of caustic potash, which, together with the fact that it refuses to strike a blue color with solution of ferric chloride, distinguishes it from morphine.

Meconine is an indifferent crystalline body obtained from opium; its solution in water is neutral.

Conine.—Formula, $C_8H_{15}N$.

Synonyms.—Conia. Coniin.

This alkaloid is volatile and obtained by distillation of crushed conium seed with dilute milk of lime. It is an oily liquid of a disagreeable tobacco odor, and a nauseating taste. It is more soluble in cold than in warm water, and its solution is strongly alkaline. Alcohol, ether and oils dissolve it freely. Conine is but little used in medicine.

Curarine.—This is a poisonous principle of the South American arrow poison *curare*. It is colorless, crystalline and forms salts with acids. This alkaloid is seldom used in medicine. It deserves further investigation.

Delphinin.—Formula, $C_{24}H_{35}NO_2$.

This exists in the seed of delphinum staphisagria. It is slightly soluble in water, but dissolves in alcohol and in ether. It is white and very acrid to the taste.

Emetine.—Formula, $C_{30}H_{44}N_2O_8$.

Synonyms.—Emetia. Emetin.

This alkaloid is a powerful emetic, and is obtained from ipecac root. It is a white powder, the chloride forming silky needles. Emetine is insoluble in ether, and soluble in alcohol.

Gelsemine (Gelseminia)

This is a poisonous alkaloid, obtained in very small amount from the root of *Gelsemium sempervirens.* It is white, bitter, and strongly basic, forming salts which are soluble and bitter. Gelsemine is slightly soluble in water, freely in chloroform or ether. This alkaloid must not be confused with the so-called resinoid (concentration) *gelsemin*, which in reality is a dried solid extract.

Hydrastine (hydrastia)—Formula, $C_{22}NH_{23}O_6$.

This is the white alkaloid of Hydrastis Canadensis, and is obtained from the mother liquor after crystallization of *berberin.* It is unfortunate that the name which is applied to this alkaloid, and which it properly deserves, as it is the characteristic alkaloid of Hydrastis C., should be also applied to the yellow alkaloid *berberin,* when it is obtained from Hydrastis. Hydrastine is white, insoluble in water, but very soluble in hot alcohol from which solution it crystallizes beautifully. It is a strong base, forming (excepting the carbazotate) salts which are soluble in all proportions of water; they refuse to crystallize. Ammonia, and the alkalies, precipitate this alkaloid from aqueous solutions of the salts. When hydrastine is chewed it imparts an acrid sensation to the fauces, and the solutions of the salts of hydrastine are very acrid and disagreeable to the taste, but not bitter.

Hyoscyamine.—Formula, $C_{15}H_{23}NO_3$.

Synonyms.—Hyoscyamia. Hyoscyamin.

This alkaloid is found in hyoscyamus, being prepared from the leaves or fruit of Hyoscyamus niger. Experiments teach the author that it is useless to endeavor to prepare hyoscyamine from even an ordinary working batch

of the crude drug, therefore, this work will not be burdened with the practical process. Hyoscyamine is a waxy mass, soluble in ether and in water. It is a strong organic base, yielding crystallizable salts with acids. Its solution precipitates white with corrosive sublimate solution. Hyoscyamine is an active poison, the treatment consists of the free use of emetics and castor oil.

Morphine.—Formula, $C_{17}H_{19}NO_3.H_2O$.

Synonyms.—Morphia. Morphin.

This the valuable alkaloid of opium. It is made by a few manufacturers upon a large scale, never being prepared in small quantities. Morphine is a white powder, insoluble in ether, slightly soluble in water, and soluble in alcohol. From its solutions by spontaneous evaporation, colorless needle-like crystals are obtained. Morphine is inodorous and very bitter. It unites freely with most acids, neutralizing them completely, and forming crystallizable salts which as a rule are soluble in water. All such are very bitter.

Tests for Impurities.—Mineral salts are shown by an insoluble residue remaining after treatment of the morphine with boiling alcohol. Also by an ash remaining after incinderation of the morphine on platinum foil.

Narcotine and Codine are shown by treating the alkaloid with ether, filtering and evaporating the filtrate upon a plate of glass. Morphine is insoluble in ether and a crystalline residue after the foregoing, indicates that the above named impurities are present.†

Cinchona alkaloids are shown by making a dilute solution of morphine in acidulated water, and then adding bicarbonate of potassium solution until in excess. *Cinchona alkaloids* precipitate, morphine remains in solution. Again, if morphine be treated with solution of caustic soda it will dissolve; cinchona alkaloids are insoluble.

Tests for Morphine.—If morphine, or a salt of morphine, be added to dilute solution of neutral ferric chloride, a blue color results. Morphine dissolves colorless

† Unless the morphine has been intentionally adulterated with the alkaloids of Cinchona, or other alkaloids which dissolve in ether.

in strong sulphuric acid; the addition of a drop of nitric acid changes this solution to purple. If the solution of morphine in sulphuric acid be mixed with a drop of solution of potassium bichromate, a green color results. Nitric acid dissolves morphine with production of an orange-red color. Dissolve eight grains of molybdate of ammonium in two fluid drachms of sulphuric acid; a drop of this solution, if rubbed with morphine or its salts, will strike a reddish-purple or a crimson color, and this will pass to greenish and finally assume a rich blue. Morphine and the salts of morphine are strong narcotic poisons.

Treatment.—Where an overdose of morphine or its salts, or of opium, has been taken, use the stomach pump at once, and thoroughly. Use emetics if the pump is not at hand; then give strong decoctions of coffee or tea. Emetics often fail to take effect in cases of opium poisoning. Apply cold affusions to the head, chest and spine, and keep the patient constantly in motion until the effects of the drug have passed, even if to do so requires rather severe whipping. Electro-magnetic shocks applied to the head and spine are valuable. Atropia or belledonna is said to antidote the effects of morphine, and repeated, small doses of either of these may be given directly or hypodermically. It should be remembered that these latter substances are also poisons and must be used with caution.

Morphine Sulphate, or sulphate of morphia, is the preparation of morphine commonly used in this country. This salt is soluble in about two parts of water, and dissolves more freely in water acidulated with sulphuric acid. It is less soluble in alcohol, and is insoluble in ether. In this country the use of this salt has become so universal as to lead dealers to supply sulphate of morphine when the term *morphine* is used in an order.

Morphine unites with other acids with formation of the respective salts. Of these the hydrochlorate (muriate) of morphine, and the acetate of morphine, are employed to a limited extent. All the salts of morphine are poisonous in overdoses, and the treatment in such cases is the same as that for poisoning by morphine.

Apomorphine.—Formula, $C_{17}H_{17}NO_2$.

Synonym.—Apomorphia.

When hydrochlorate of morphine is sealed in a strong vessel with an excess of hydrochloric acid, and a temperature of 300° F. is applied for some hours, a molecule of water is abstracted from the molecule of morphine. The resulting salt has the composition ascribed above with the addition of a molecule of hydrochloric acid. The base (apomorphine) may be separated by means of acid carbonate of sodium, and then dissolved in chloroform, from which it is obtained by evaporation. Apomorphine differs from morphine in its physiological action, although differing in composition only by a molecule of water. Apomorphine is a violent *emetic*, and one-tenth of a grain, or less, injected hypodermically, will produce vomiting in a few moments.

Narcotine.—Formula, $C_{22}H_{23}NO_7$.

Synonyms.—Narcotin. Narcotia.

This alkaloid remains in opium after it (opium) is extracted with water, and from which it may be dissolved with acidulated water. It is white, crystalline, and insoluble in cold water. Alcohol dissolves it to an extent, chloroform dissolves it freely. It does not change solution of ferric chloride to blue (distinction from morphine). It has no soporific influence. Narcotine must not be confused with—

Nicotine (nicotia), the volatile alkaloid of tobacco. Nicotine is soluble in water and in ether. It has a disagreeable, nauseous odor, and a burning taste. Chlorine strikes with it a blood red color, and tannic acid gives a white precipitate. Nicotine is a deadly poison, and remarkably quick in its action. There is no chemical antidote and the treatment seems simply to consist in the free use of emetics (if the drug fails to vomit) or the stomach pump. Afterward administer oils and mucilaginous drinks.

Physostigmine.

Synonyms.—Physostigmia. Eserina.

This is the active principle of Calabar bean (Physostigma veneosum), and is made from extract of Calabar. It is colorless, soluble freely in alcohol, ether, and chloroform, and is less soluble in water. Physostigmine is tasteless, and unites with acids to form tasteless salts which as a rule are soluble. Alkalies decompose it with formation of a red color. The solutions of physostigmine and its salts yield precipitates with the alkaloidal reagents, such as mercuric chloride, chloride of gold, etc.

Physostigmine is very poisonous, and the treatment consists in the use of emetics or the stomach pump, followed by stimulants. Inasmuch as Belladonna antagonizes the effect of Calabar bean upon the pupil of the eye, each has been recommended as an antidote for the other.

Piperine.—Formula, $C_{17}H_{19}NO_3$.

Synonym.—Piperin.

This is a weak alkaloid, and is derived from black pepper. It is white, and tasteless if pure, but often has a yellow color, and a peppery taste. It is neutral, soluble in alcohol and ether, and dissolves more freely in cold than in hot water. It dissolves freely in acetic acid.

Sanguinarine.—Formula, $C_{19}H_{17}NO_4$.

Synonyms.—Sanguinarina. Chelerythrin. Sanguinarin.

This alkaloid was originally discovered in the root of *Chelidonium majus*, and named chelerythein. The fact that it is yielded most abundantly by the rhizome of *Sanguinaria Canadensis*, and that considerable amounts of its salts from this source are used in medicine, always under the name *sanguinarine* or *sanguinarina*, we are led to give preference to the name *sanguinarine*. This alkaloid is made by evaporating an alcoholic extract of Sanguinaria Canadensis to dryness, and extracting the residue with water. This is filtered, and to the filtrate, ammonia water in a slight excess is added. The bluish precipitate is then separated by filtration and dried. The dry mass is

cautiously powdered, and exhausted with ether, and the mixture filtered. The filtrate is agitated with a little nitric acid (not in excess) whereby *nitrate of sanguinarine* is produced, and this being almost insoluble in ether, precipitates as a yellowish-red magma, which becomes crimson when dried. This powder, if triturated with an ethereal solution of ammonia gas, gives the alkaloid in a very pure form to the ether, and from this solution it may be obtained by evaporation. Sanguinarine is a white powder, insoluble in water, soluble in alcohol and in ether. It is intensely irritating when inhaled, even a minute portion of dust rendering the air of a room almost irrespirable. Sanguinarine is remarkable for the fact that all of its salts are *red*, the alkaloid being white. It is a strong base and forms salts with even the weak organic acids. These are made by simply neutralizing the alkaloid with a weak solution of the necessary acid, and evaporating to dryness. Of the salts, nitrate of sanguinarine and sulphate of sanguinarine are in considerable demand.

Strychnine.—Formula, $C_{21}H_{22}N_2O_2$.

Synonyms—Strychnia. Strychnin.

This active poison is an alkaloid of Nux vomica. It is white, odorless and intensely bitter. It dissolves in 6000 parts of cold water and even in this proportion imparts a distinct bitterness. Hot alcohol dissolves one-tenth part of strychnine, and from this solution the alkaloid crystallizes in prisms upon cooling. Ether and alkaline solutions can not dissolve strychnine. Strychnine is a strong base, and unites with acids to form salts, mostly soluble in water. Owing to the fact that strychnine is in common use as a medicine, and that its poisonous properties are well known, cases of poisoning by this drug are rather frequent.

Tests for Strychnine.—Mix a little strychnine with a little concentrated sulphuric acid and add oxide of cerium. A beautiful blue color results, which changes slowly to purple. Peroxide of lead produces a blue color when it is added to the solution of strychnine in sulphuric acid.

If to a solution of strychnine in sulphuric acid, a little nitric acid be added, and then binoxide of manganese, a purplish-violet color results. The foregoing tests are characteristic, and very delicate.

Treatment of poisoning where strychnine, or salts of strychnine have been taken. If the mouth can be opened, empty the stomach by means of a stomach pump or emetics. Keep the patient as quiet as possible, and let the attendants move cautiously and speak low. A loud noise or a jar will often throw the person into convulsions. Keep the room dark. Chloral hydrate or chloroform may be administered until the patient is influenced. Resort to artificial respiration if necessary.

Veratrine.—Formula, $C_{32}H_{52}N_2O_8$.

Synonyms.—Veratrin. Veratria.

This alkaloid is obtained from *Veratrum album.* It is seldom employed in medicine in consequence of its violent action. Veratrine is white, and soluble in alcohol. It is acrid and bitter to the taste. If the dust of this alkaloid be inhaled, even in minute proportion, it produces violent sneezing, followed by subsequent irritation of the mucous membrane.

At one period it was thought that *Veratrum viride* contained this alkaloid, but further investigation demonstrated that such was not the case.

GLUCOSIDES.

This class of organic bodies yields glucose as a product of decomposition when boiled with acidulated water. They are complex in composition, odorless, usually neutral, but sometimes slightly acid in reaction. As a rule they dissolve in hot alcohol. Only a few glucosides are employed in a pure form as medicines. Many are known to exist as active principles of plants, but which are not purified. If a glucoside be boiled with dilute sulphuric acid, and the solution be then neutralized with carbonate of barium and filtered, the filtrate will give, with Fehling's solution, the reaction which characterizes glucose.

Glucosides are sometimes called neutral or indifferent principles. We use the syllable *in* as a termination instead of *ine* as with alkaloids.

Salicin.—Formula, $C_{13}H_{18}O_7$.

Salicin is derived from willow bark (Salix), and from thence obtained its name. It is also found in poplar bark. Salicin is white, crystallizable, odorless, and very bitter. It dissolves in half its weight of boiling water, and in three parts of boiling alcohol. Sulphuric acid produces with it a purple-red color. If boiled with dilute sulphuric acid, glucose and a body called *saligenin* results. Saligenin in aqueous solution gives a deep indigo-blue color with ferric salts. Oxidizing agents, such as chromic acid, convert it into *salicylic acid* and salicylic aldehyd. From this fact the name "salicylic acid" was derived.

Santonin.—Formula, $C_{15}H_{18}O_3$.

Synonyms.—Santoninum. Santonic acid.

This substance is obtained from wormseed. It is now accepted to be the anhydrid of a weak acid and is called *santonic anhydrid.* Santonin is white, crystalline, odorless, and slightly bitter to the taste. It turns yellow by exposure to the light, dissolves in three parts of boiling alcohol, and in forty-three parts of cold alcohol, in five-thousand parts of cold water, and in about four and one-half parts of cold chloroform. Santonin unites with alkalies to form soluble salts.

Adulteration and Tests.—Santonin is often adulterated when purchased from irresponsible parties. We once noticed a sample containing 75 per cent. boric acid. Boric acid will not dissolve in chloroform, and imparts a green color to an alcohol flame. If a specimen of santonin refuses to dissolve perfectly in an excess of chloroform, the insoluble matter is adulteration, and is likely to be boric acid, or potassium chlorate. If santonin dissolves to any appreciable amount in cold water, soluble impurities are present. If it is yellow, and *very* bitter, picric acid is likely present, and the aqueous solution will then stain the nails yellow, and form an insoluble precipitate with diluted solution of berberine.

Cathartic Acid is said to be the cathartic principle of senna. The high atomic composition $C_{180}H_{192}N_4O_{82}S$, is ascribed to it. This substance is not used in the pure state.

Amygdalin, $C_{20}H_{27}NO_{11}$, is a glucoside of bitter almonds and which with emulsin yields hydrocyanic acid. The chemical history of this glucoside is interesting to theorists, but hardly appropriate for consideration in a work of this description. It is decomposed by both acids and alkalies.

Digitalin, $C_{27}H_{45}O_{15}$, is prepared by a tedious operation from *Digitalis purpurea.* It is an active poison, has a yellowish color, is neutral, odorless and very bitter. In this country digitalin is not used in medicine, but it is officinal in the British Pharmacopœia.

Glycyrrhizin.—This is the sweet principle of liquorice root (Radix Glycyrrhizæ). It unites with ammonia and with quinine, forming with the latter body an insoluble compound. It is mainly owing to this fact that solutions of liquorice are employed to disguise the taste of quinine. In such cases it must be remembered that the quinine *precipitates* and is not in the solution. Mixtures of this description must be well shaken before each dose is administered.

Colocynthin is the active principle of colocynth, and a glucoside. It has the composition, $C_{56}H_{84}O_{23}$, and is soluble in water and in alcohol, but insoluble in ether. Sulphuric acid strikes with it a red color. This glucoside must not be confused with the so-called "resinoid," *Colocinthin,* which is not a definite compound.

Æsculin, is a white glucoside of the horse chestnut (Æsculus Hippocastanum), and recently asserted to be identical with gelseminic acid. It is white, crystallizable, and strikes a blue fluorescent color with alkaline aqueous solutions.

Jalapin is the active principle of scammony. The article sold in this country as *Jalapin* under the elastic

name "concentration," or "resinoid," is either a dried extract of jalap, or the resin of jalap.

Colvolvulin is a glucoside and constitutes the larger part of resin of jalap (see jalap resin.) There are many glucosides besides those we have named, but such are not employed in medicine. Saponin, Smilacin, Pectin, etc., are examples.

RESINS.

This is a class of bodies very widely distributed throughout the vegetable kingdom. True resins seem to be formed from the oxidation of volatile oils. They are *insoluble in water.* Resins often contain admixed volatile oils, which impart an odor. Alcohol (pure) dissolves all resins. Ether and oils dissolve most resins. Some resins are tasteless, others bitter. Some are transparent and colorless, others colored or even black. All resins contain carbon and hydrogen, many contain carbon hydrogen and oxygen; none contain nitrogen. Solution of the caustic alkalies dissolves most resins, soapy mixtures resulting. Resins fuse by heat, then decompose and burn with smoky flames. If most resins be submitted to dry distillation oily products result. These are generally hydrocarbons. Many of the resins of pharmacy are not true resins, but approach them more nearly than they do any other class of organic bodies. Others are now proven to be mixtures of substances, but long usage retains for such the name resin, and we shall consider them among the resins †

Resin.

Synonyms.—Rosin. Colophony. Resina. Resinæ.

This is typical of the resins, and is the most common example. It is made by distilling crude turpentine, obtained from the pine, whereby the volatile oil is driven from the retort, resin remaining. Resin is found in

† Some consider, under the name "resinoids," all bodies which partake of the characters of resins. In this country the term "resinoid" has been applied by some indiscriminately to a class of bodies, embracing alkaloids, extracts, etc., and often of a secrete nature.

market ranging in color from a deep brown to a very light straw color. The lighter colors are most valuable. Resin consists mostly of an acid called sylvic or abietic acid ($C_{20}H_{30}O_2$), which may be obtained by treating resin with alcohol, filtering, and adding water to the filtrate. Minute crystals of the acid are thereby precipitated. These may be obtained in large crystals by dissolving in hot alcohol and afterward cooling the solution. Resin is of great use in pharmacy, entering into plasters, cerates and ointments.

Resin of Jalap.

Synonyms.—Resina. Jalapa.

This is a resinous substance and is obtained from jalap. It is made by exhausting the powdered tubers with alcohol, distilling the alcohol, then pouring the residue into cold water. The resinous matter separates, and by repeated washing is freed from glucose and the other substances which are soluble in water. Resin of jalap is a mixture of two resins, one soluble in alcohol and in ether, the other soluble in alcohol, but insoluble in ether. This latter is known generally as *convolvulin*, although Kayser named it *rhodearetin*, in 1844. Convolvulin is almost colorless when pure and is odorless and tasteless. It is said to have the composition $C_{62}H_{100}O_{32}$, and to be the anhydrid of *convolvulic acid*, $C_{26}H_{106}O_{35}$. When it is boiled with dilute acids it splits into glucose and convolvulinol, $C_{26}H_{48}O_6$. Convolvulin is soluble in alkaline solutions, but is not reprecipitated when they are rendered acid. The portion of jalap resin which is soluble in ether, is soluble in alkaline solutions and reprecipitated from them by acids. The reader must bear in mind that *jalapin* is the active principle (resin) of scammony, and that *convolvulin* is the active principle of jalap. It would be better were these names reversed and consistent. The so-called resinoid *jalapin* is simply purified resin of jalap dried and powdered. Owing to the moderate price of jalap, adulteration of resin of jalap is hardly an object at the present. The most likely admixture is powdered jalap, and this can be shown by treating a specimen of

jalap resin with an excess of boiling alcohol. The woody matter will in this case remain as a residue.

Resin of Scammony.

Synonym.—Resina scammonii.

This resinous body is made from powdered scammony (scammony root B. P.) in the same manner that jalap resin is produced from jalap, boiling alcohol being used in place of the cold alcohol of that operation. Resin of scammony is of a brown color and is perfectly soluble in alcohol and in ether. It seems that the larger part of this resin is a glucoside of the composition, $C_{34}H_{56}O_{16}$, and is unfortunately named *jalapin*. Resin of scammony turns red when treated with sulphuric acid. It dissolves in solution of caustic alkalies, and an excess of the mineral acids will not precipitate such a solution.

Dragon's Blood.

Synonym.—Resina draconis.

This is an impure resin of a dark brownish-red color externally, brighter red internally. It is found in market in cylindrical sticks, and is the product of the fruit of a palm, Calamus draco (*Wild*). Dragons blood consists mainly of a resin soluble in alcohol, but which is not used in medicine in a purified condition.

Resin of Podophyllum.

Synonyms.—Podophyllin. Resina podophylli.

When the rhizomæ of *podophyllum peltatum* are powdered and digested in alcohol this so-called resin is extracted. If the alcohol be evaporated from such a tincture, and the residue poured into cold water, the resinous substance settles, and may be purified from admixtures by repeated washings with cold water. The name *podophyllin* was originally ascribed to this substance. Resin of podophyllum is mostly soluble in alcohol, but after each solution of it and the subsequent exposure in drying, a change results which renders a certain amount insoluble in fresh portions of alcohol. About one-fourth of the resin dissolves in ether, and this is said to be more strongly cathartic than the other portion. Resin of

podophyllum differs in color according to the mode of precipitation and subsequent drying. It can not be made of a yellow, or a greenish yellow color without the use of chemicals. That article which is sold on the market of a yellow color is made by precipitating the resin by means of alum water. To an extent resin of podophyllum dissolves in hot water.

Resin of Leptandra.

Synonym.—Leptandrin.

This is a product of the dried root of *Leptandra Virginica*. It is made according to the process for making resin of podophyllum. Resin of leptandra is black, very much resembling asphaltum and, has a peculiar odor. At one time this article was thrown on the market (leptandrin) as the active principle of Leptandra Virginica, and was included among the so-called "resinoids." The experiments of the writer lead him to believe that this resinous body is produced by the decomposition of a glucoside, and that it does not exist in the sound portions of green Leptandra root.

Resin of Cimicifuga.

Synonyms.—Cimicifugin. Macrotin.

This body is made by evaporating the alcoholic tincture of Cimicifuga root to the consistence of an extract, and then pouring it into water. The resinous substance which separates is then washed with water and dried by exposure to the air. This has a brownish color and a peculiar smoky odor. It imparts a sweetish taste, if chewed. Considerable amounts of this resinous body are sold under the terms given as synonyms at the head of our article. It was originally introduced as a resinoid, and at present the position is appropriate.

OLEORESINS.

These are mixtures, mainly of resinous bodies and volatile oils. They are more or less soft, and are obtained by evaporating an ethereal or an alcoholic tincture to the consistence of a syrup, and then pouring the residue into

cold water. The mixture of oil and resin now settles. The oleoresins often depend more or less upon the presence of other substances for their medicinal values, the actual oleoresin serving to envelop such bodies and prevent their solution in the water. We do not feel that it is necessary to dwell at length upon the oleoresins of commerce as the subject is properly connected with pharmacy. They may be named as follows:—

OLEORESINA.	OBTAINED FROM
Aletrin	Aletris farinosa.
Apocynin	Apocynum cannabinum.
Asclepedin	Asclepias tuburosa.
Capsicin or Capsici	Capsicum.
Cubebin or Cubebæ	Cubebs.
Felicis	Male fern.
Helonin	Chamælirium luteum.
Irisin	Iris versicolor.
Lupulinæ	Lupulin.
Piperis	Black pepper.
Xanthoxylin	Xanthoxylum fraxineum.
Zingiberis	Zingiber.

RESINOIDS.

Under this heading some years ago, certain manufacturers of pharmaceutical preparations classed a number of substances. To the majority of these the name "resinoid" could not be given with any degree of propriety. Many of the so-called resinoids were true alkaloids, some were resins, others were dried solid extracts, or bodies of a like nature. We have classed among the resins or the alkaloids all of those which we believe deserving of notice. If an injustice has been shown any of the so-called "resinoids" the claims of such articles will be cheerfully considered if any person will notify the author of that fact. It might, perhaps, be justly argued that resin of podophyllum is not a true resin, and should be classed as a resinoid.

VOLATILE OILS.

These are obtained by distillation of plants with water. Volatile oils are also known as *essential* or *ethereal* oils. These bodies are usually mixtures of *elœoptens* † and *stearoptens*,‡ and are prone to absorb oxygen, so that age alters their properties to a great extent. Numbers of volatile oils are composed of two or more isomeric elœoptens mixed with stearoptens. The taste of the volatile oils is usually hot and burning. When freshly distilled the majority of them are colorless, but many darken by age. When some volatile oils are submitted to fractional distillation, they yield several isomeric elœoptens, differing mainly in their boiling.

Volatile oils are subject, in certain cases, to great adulteration. The ingenuity displayed by those engaged in the business of adulterating them would be admired if brought forward in a better cause. No sooner is one adulterant exposed than another takes its place. Owing to the variation in composition, gravity, and other properties of many oils by age and exposure, even if pure, it is difficult to give fixed rules for detecting impurities. This fact, and also the fact that many oils are isomerous, forces us to rely more upon the honor of manufacturers and dealers than upon tests for impurities. The writer has been led to the opinion that adulterated oils, as a rule, are obtained by such as will not pay the market value of pure oils, and that in such cases the adulteration is in proportion to the price at which the oil is demanded. Of course, there are exceptions; it may be that certain imported oils are always adulterated, but it is certain that the volatile oils made in this country may be purchased

† *Elœoptens* are hydrocarbons, usually isomeric with oil of turpentine, $C_{10}H_{16}$. They are liquid, and often hold in solution stearoptens. By absorption of oxygen they may change to stearoptens.

‡ *Stearoptens*. These are the result of the oxidation of elœoptens, having the composition $C_{10}H_{16}O$. They are solid or camphor-like bodies, and often crystallize from volatile oils that are exposed to the air, or which are reduced in temperature.

perfectly pure, if the purchaser is willing to pay their real value.

Oil of Anise.

Synonym—Oleum anisi.

This is the distilled oil from the fruits of *Pimpinella anisum* and *Illicium anisatum*. That from the latter is properly oil of staranise, but the two are not distinguished in commerce, and are sold under the general name, oil of anise. Oil of anise is almost colorless or of a pale straw color when fresh, but darkens by age. The specific gravity is from 0.97 to 0.98, but according to Zeller it may rise in old oils to 1.028. Anise oil solidifies at a moderately high temperature, forming a crystalline mass. This varies in accordance with the origin of the oil, and subsequent age and exposure. The true anise oil, if fresh, is said to congeal at about 10° C. (50° F.) while oil of staranise requires a temperature of 2° C. (35.6° F.)

Oil of anise is composed mainly of a body named *anethol*, $C_{10}H_{12}O$, which exists in two isomeric forms, one solid at ordinary temperatures, the other liquid. Anise oil also contains a small amount of an elœopten. If anise oil be adulterated with wax or with spermaceti, these bodies will remain insoluble upon digesting the mixture with eighty per cent. alcohol. Alcohol is shown by agitating the liquefied oil with water, the decrease in bulk of the oil being due to loss of alcohol.

Oil of Artemisia.

Synonyms.—Oil of wormwood. Oil of absinthium.

This is obtained by distillation of the herb *Artemisia absinthium*, and when crude has a greenish color. If redistilled, however, it may be obtained nearly colorless. Oil of wormwood possesses the odor of the herb, has a pungent taste, but is destitute of bitterness.

Oil of Bergamot.

Synonym.—Oleum Bergamii.

This oil is yielded by the fruit of *Citrus limetta*, D. C., and is imported from Italy. It is used only as a perfume.

The color is greenish, the specific gravity from 0.86 to 0.88; the odor is pleasant to most persons. Bergamot oil dissolves in its own weight of alcohol. The fresh oil contains a waxy matter ($C_9H_6O_3$) which crystallizes upon standing, and is known as *bergamot camphor*. The principal part of the oil is a mixture of elœoptens of the formula, $C_{10}H_{16}$. Bergamot oil is adulterated to suit the price which consumers desire to pay for it. The adulterants are turpentine oil, alcohol and other and cheaper volatile oils, although occasionally a fixed oil is introduced. The best test is officinal alcohol, in which bergamot oil should dissolve freely, bulk for bulk This, of course, will not detect alcohol. If alcohol be present the fact may be shown by shaking together measured amounts of the suspected oil and water. Upon standing the oil will separate, and the increase in bulk of the water will give about the proportion of alcohol present in the bergamot oil.

Oil of Capsicum.

Under this name an article is sold which in reality is a concentrated solution of the principles upon which capsicum depends for its medicinal value. This is made by exhausting capsicum with stronger alcohol, evaporating the alcohol until the residue is of the consistence of a syrupy extract, then digesting the extract in twenty times its bulk of sulphuric ether. The ethereal solution is then filtered and the filtrate evaporated to an oily consistence.

This so-called oil of capsicum is intensely fiery and represents in a concentrated degree the properties of capsicum. It has a red color, and if old, is disposed to deposit a brown sediment due to the decomposition of the red astringent principles. It contains *capsicine*, and such other of the constituents of capsicum as are soluble in ether, these being held in liquid form by the natural fixed oil of the capsicum.

Oil of Caraway.

Synonyms.—Oil of carum. Oleum carvi. Oleum cari.

This is the volatile oil of the fruit of *Carum carvi*, and is obtained in the usual manner. It is colorless and mo-

bile when fresh, but becomes colored and gummy by age and by exposure. It has the agreeable odor and taste of caraway seed. Oil of caraway dissolves in all proportions in alcohol. It is liable to be adulterated with turpentine oil, and this can be carried to a great extent without altering the color and taste of the caraway.

Oil of caraway is a turpene, $C_{10}H_{16}$ (carven), and this by oxidation forms *carvol*, $C_{10}H_{16}O$. Carven distills at 175°, carvol between 175° F. and 230° F.

Oil of Chenopodium anthelminticum.

Synonyms—Oleum Chenopodii. Oil of wormseed.

This oil is obtained by distillation with water of the fruit of Chenopodium anthelminticum. It is colorless when fresh, but changes to brown, and thickens by age and exposure. The taste and odor are that of wormseed. This oil is composed of an elœopten, $C_{10}H_{16}$ and a stearopten, $C_{10}H_{16}O$. It is often adulterated with oil of turpentine.

Oil of Cinnamon.

Synonym.—Oleum Cinnamonii.

This, according to the U. S. P., is the distilled oil of Ceylon cinnamon. That in common use is oil of cinnamon cassia, or oil of cassia. These oils resemble each other so nearly that in reality little preference can be made between them. The finer qualities of cassia oil approach, if not equaling, in color and odor the oil of Ceylon cinnamon. Cinnamon oil is pleasant to most persons, resembling in odor and taste the ordinary cinnamon bark. Oil of cinnamon and oil of cassia have a light yellowish color when recent, which changes to red or even brown by age and exposure. Oil of Ceylon cinnamon is said to have the specific gravity of 1.035 or 1.055, while the specific gravity of oil of cassia is from 1.060 to 1.065. Specific gravity alone, however, is of little value as a detectant of admixtures, for artful adulterators can produce the desired gravity without trouble.

Oil of cinnamon consists chiefly of *cinnamic aldehyd*, C_9H_7OH. By age *cinnamic acid*, $C_9H_8O_2$, is formed to an extent, accompanied by a reddish resin. Cinnamon oil should mix without residue with its own bulk of officinal alcohol.

Oil of Cloves.

Synonym.—Oleum caryophylli.

This volatile oil is made by aqueous distillation of cloves. Oil of cloves is colorless or pale yellow when recent, but becomes brown by age. It has the odor and taste of cloves, but very much increased. It dissolves readily in alcohol, has a specific gravity of from 1.034 to 1.056, and boils at about 240° C. (455° F.) Oil of cloves contains *eugenic acid*, $HC_{10}H_{11}O_2$, a hydrocarbon of the composition, $C_{10}H_{16}$, and a stearopten.

Oil of Copaiba,

Synonym.—Oleum copaibæ.

The essential oil is obtained by distilling copaiba balsam with water. It is colorless, and thin when fresh, but becomes thicker and oily by age and exposure. It has the disagreeable odor and taste of copaiba, specific gravity of from 0.88 to 0.91, and consists of isomerous elœoptens of the composition, $C_{10}H_{16}$, or of elœoptens with a turpene polymeric with oil of turpentine, and in this case is represented by $C_{20}H_{32}$.

Oil of Coriander.

Synonym.—Oleum Coriandri.

This essential oil is derived from the fruit of Coriandium sativum (*Lin.*), and in the usual manner for making essential oils. It is clear, has the odor of the fruit from whence it is derived, and the specific gravity of from 0.86 to 0.87. When this oil is distilled from phosphoric anhydrid it yields an isomerid of $C_{10}H_{16}$ known as *camphene*. The oil of coriander is supposed to be $C_{10}H_{16}$-H_2O. It reacts violently if mixed with iodine, often exploding. The adulterant is turpentine oil.

Oil of Cubebs.

Synonym.—Oleum cubebæ.

This is made by distilling Cubeb berries with water. It is colorless, has the odor of cubebs, an aromatic taste and the specific gravity 0.92. It is mainly composed of *cubebene,* or $C_{15}H_{24}$,

Oil of Erigeron.

Synonym.—Oleum Erigerontis Canadensis.

Oil of erigeron is made by distillation with water of the herb of Erigeron Canadense. (*Lin.*) It is colorless when fresh and has a rather disagreeable odor. It becomes red by age. Oil of erigeron is sometimes sold for *oil of fireweed,* but the true oil of fireweed is derived from the *Erechthites hieracifolius.*

Oil of Fennel.

Synonym.—Oleum Fœniculi.

This is obtained by distilling bruised fennel-seed with water. It has the odor and taste of fennel, the specific gravity 0.968, and consists of liquid and solid *anethol,* and a liquid hydrocarbon of the composition, $C_{10}H_{16}$. Oil of fennel mixes with its bulk of alcohol without precipitation.

Oil of Juniper.

Synonym.—Oleum Juniperi.

This oil is obtained from juniper berries by distillation with salt water. It has the odor of the berries, is colorless, and consists of a mixture of hydrocarbons of the general composition, $C_{10}H_{16}$. By combination with water a hydrate is formed, $C_{10}H_{16}H_2O$. Oil of ripe juniper berries boils at 205° C. The oil which is derived from unripe berries contains in addition an oil which boils at 155° C. Oil of juniper wood is sometimes employed as a liniment. It is colorless and thin when fresh, but darkens and becomes thick by age.

Oil of Cade is a tar-like oil obtained by destructive distillation of the wood of Juniperus oxycedrus. Oil of juniper berries and wood is largely adulterated with turpentine oil.

Oil of Lavender.

Synonym.—Oleum Lavandulæ.

This is obtained by aqueous distillation from the flowers, or the flowering tops of Lavandula vera, D. C. It is colorless and limpid when fresh, becoming thick and red by age. It is fragrant, has the specific gravity of from 0.87 to 0.91 and mixes with its own weight of alcohol. Oil of lavender contains a hydrocarbon of the composition, $C_{10}H_{16}$, and a stearopten which (Dumas) resembles common camphor. Turpentine oil is employed as the adulterant of oil of lavender, and such mixtures are less soluble in alcohol than the pure oil. In commerce inferior oils, derived from different parts of the plant, are sold simply as *oil of lavender*. The best quality, that derived from the flowers, is known as *oil of garden lavender*. This last should be used by pharmacists.

Oil of Mustard, (Volatile).

Synonym.—Oleum sinapis.

This oil is produced when black mustard seed is deprived of its fixed oil by pressure, the residuum being distilled with water. It does not exist in the seed, but is the result of the decomposition of *myronic acid*, or *myronate of potassium*, under the influence of a ferment named *myrosin*. Oil of mustard has the composition, $C_3H_5.CNS$. It is in reality, sulphocyanide of allyl. It is somewhat soluble in water, is colorless or nearly so, if pure, and has a powerful and pungent odor. This is so penetrating as to take the breath from a person who chances to incautiously inhale it. The writer witnessed a case in which a boy after inhaling it dropped, at once, insensible. Fresh air and cold water, the latter applied to the head and back, restored him to consciousness. This oil should be handled cautiously, and never confounded with the expressed (fixed) oil of mustard, which is bland and inert.

Oil of Lemon.

Synonym.—Oleum limonis.

This is the volatile oil of lemon peel, and is obtained by breaking the oil cells of the rind and draining from

them the oil. Inferior oil of lemon is made by distillation with water. Oil of lemon has the specific gravity of about 0.852, is yellowish and should possess the fragrance of fresh lemon peel without contamination with a turpentine odor. Commercial oil of lemon contains usually some white flocculent matter which separates by standing. The clear oil should dissolve perfectly in eight parts of alcohol, and yield a clear solution. By age oil of lemon undergoes a change, acquiring a turpentine-like odor. This is retarded by the addition of about five per cent. of alcohol. Two hydrocarbons of the composition, $C_{10}H_{16}$, constitute the bulk of lemon oil. These absorb oxygen by exposure and give rise to a precipitate called *citroptene*. Oil of lemon is extensively adulterated, the adulterant usually being turpentine oil.

Oil of Lobelia.

The seed of Lobelia inflata contains a large amount of a fixed bland oil, and which must not be confused with the article sold as "oil of lobelia." This is made by exhausting powdered lobelia seed by means of stronger alcohol, then evaporating the alcohol and extracting the residue with twenty times its bulk of ether. The ethereal solution is to be filtered and then evaporated until all of the ether is expelled. The result is the so-called "oil of lobelia." It is in reality the fixed oil of lobelia containing in solution the active principles of the seed. It represents the emetic properties of the seed in a very much increased degree. Such properties do not result from the oil, but from the substances the oil holds in solution.

Oil of Nutmeg.

Synonyms—Oleum myristicæ. Volatile oil of nutmeg.

Oil of nutmeg is obtained by aqueous distillation of ground nutmegs. It is yellowish, limpid, possesses the odor of nutmegs, and has the specific gravity of from 0.92 to 0.95. It consists mainly of a hydrocarbon, $C_{10}H_{16}$, and by absorption of oxygen deposits a stearopten, *myristicin*. Nutmegs yield by pressure a fixed oil which is of a firm consistence and a dark color. This is mainly a fat, *myristin*, of the composition $C_3H_5 3C_{13}H_{27}O_2$.

Oil of Origanum.

Synonym.—Oleum origani.

This oil is obtained by distilling with water the fresh herb, Origanum vulgare. (*Lin.*) It is limpid and has, when fresh, a yellow color. By age it becomes brownish. The chemistry of this oil has yet to be determined with satisfaction. Oil of origanum of commerce is generally the volatile oil of Thyme (Thymus vulgaris, *Lin.*), usually at the present day destitute of thymol. The cheap commercial oil of origanum contains turpentine oil.

Oil of Pennyroyal.

Synonym.—Oleum hedeomæ.

This is the volatile oil of Hedeoma pulegioides (*Lin*), and is obtained by distillation of the herb with water. It is colorless when fresh and has the odor of pennyroyal. By age it darkens and the odor becomes like turpentine. The specific gravity is about 0.95. Alcohol dissolves it freely.

Oil of Peppermint.

Synonym.—Oleum menthæ piperitæ.

The fresh herb of Mentha piperita (*Lin.*) is distilled with water in the usual manner for obtaining volatile oils. Peppermint oil is colorless when fresh, although some specimens have a yellowish or a greenish tint. By age and exposure it becomes viscid and dark colored. It has the specific gravity of about 0.90, dissolves perfectly in four parts of officinal alcohol. Oil of peppermint contains a liquid hydrocarbon, $C_{10}H_{18}$, termed *menthene*, and a solid body known as *menthol* or mint camphor. This latter substance is the principle to which the oil owes its fragrance and taste. It has the composition, $C_{10}H_{20}O$. and seems to be a hydrate of menthene ($C_{10}H_{18}H_2O$.). The adulterants of peppermint oil are generally turpentine oil, castor oil or alcohol.

Oil of Rose.

Synonyms.—Oleum rosæ. Otto of rose.

This valuable oil is made by distilling water with the flowers of a variety of Rosa damascena. (*Miller*). Oil of

rose is employed as a perfume for certain pomades, and as an ingredient of cologne and other perfumes. It has a yellowish color. It solidifies at about 50° F., if slowly cooled to that point. Oil of rose consists of an elæopten and a stearopten. Commercial oil of rose varies in purity according to the price for which it is sold. The adulterants are olive and castor oils, alcohol, spermaceti and oil of rose geranium. The property of congelation is imparted to such mixtures by the spermaceti.

Oil of Rosemary.

Synonym—Oleum Rosemarini.

This is, when fresh, a colorless, fragrant oil, obtained by aqueous distillation from the fresh tops of Rosemarinus officinalis. (*Lin*). The specific gravity ranges form 0.88 to 0.91. It contains a stearopten which resembles camphor, and which is doubtless produced by the oxidation of hydrocarbons.

Oil of Rue.

Synonym.—Oleum rutæ.

This is obtained in the usual manner by distillation of water with the fresh herb of Ruta graveolens. (*Lin.*) It is colorless if fresh, changing to yellow by age. At a low temperature it congeals, forming shining scales. Oil of rue contains a hydrocarbon of the composition, $C_{10}H_{16}$, and an oxygenated body which constitutes the bulk of the oil. This latter substance has been proven to be a ketone, and to have the composition, $CH_3.CO.C_9H_{19}$.

Oil of Sassafras.

Synonym.—Oleum sassafras.

Oil of sassafras is made by distilling the bark of the root, or even the entire root of sassafras officinale, with water. It is colorless when pure, but changes to red by age. It has the specific gravity of about 1.090, and a pleasant taste. The bulk of sassafras oil is composed of $C_{10}H_{10}O_2$, and has been named *safrol.* Oil of sassafras also contains a hydrocarbon of the composition, $C_{10}H_{16}$. Sassafras oil is adulterated for the benefit of those who

desire a cheap article. In such cases oil of turpentine is usually employed, but this is not necessarily the case. Oil of sassafras will dissolve iodine quietly, is violently decomposed by bromine, and acquires a red color in the cold if acted upon by nitric acid.

Oil of Savin.

Synonym.—Oleum sabinæ.

This is distilled from savin (Juniperus sabina) by the usual method for obtaining volatile oils. It is colorless when fresh, has the specific gravity of about 0.90 or 0.91, and has the disagreeable odor and taste of savin. Savin oil consists of a terpene, $C_{10}H_{16}$, closely resembling turpentine oil.

Oil of Stillingia.

Synonym.—Oil of Queen's root.

The root of Stillingia sylvatica (*Lin.*) yields by treatment with ether, or with benzine, a solution of a fixed oil, This oil is afterward obtained in a pure form by evaporating the ether or benzine, dissolving the residue in alcohol, filtering this solution and evaporating the filtrate until the alcohol is expelled. Oil of stillingia made in this manner is clear and of a greenish color, about the consistency of olive oil, and like that oil congeals upon cooling. It is soluble in alcohol and in ether. When swallowed it leaves an acrid sensation in the throat. The article sold on the market under the name "oil of stillingia" is entirely a different preparation. It is simply a concentrated alcoholic extract. To make this, exhaust powdered stillingia root *well dried*, with *stronger* alcohol, and evaporate the percolate to the consistence of a thin syrup. Mix with this when cold, a like bulk of ether, and agitate. In twenty four hours filter the result, and evaporate the filtrate to the consistence of a thick syrup. "Oil of stillingia," as thus made, contains the larger part of the therapeutical principles of the root, and is free from most of the gummy and inert extractive matters. It has a red color, due to the presence of tannates, and sometimes gelatinizes in consequence of the presence of such sub-

stances. We admit that this preparation may be an excellent therapeutic agent, but we insist that it does not deserve the name "oil," and that such a title is a misnomer.

Oil of Turpentine.

Synonyms.—Turpentine. Spirit of turpentine. Oleum terebinthinæ.

Various species of Pinus exude an oleoresin which is properly *turpentine.* This, when distilled, furnishes a volatile oil, *oil of turpentine,* the residue within the retort being rosin. Turpentine is typical of the terpenes. It has the composition, $H_{10}C_{16}$, from whatever source obtained, but it differs very much in optical and other respects, according to origin. It is colorless and has a peculiar disagreeable odor. It is insoluble in water, dissolves in about ten parts of officinal alcohol, is oxidized by nitric acid and potassium permanganate, and if it is mixed with bromine or iodine a violent reaction results. It unites with hydrochloric acid to form the compound, $C_{10}H_{16}HCl$., *turpentine camphor.* The same name is applied to a combination of turpentine oil and water, $C_{10}H_{20}O$. Oil of turpentine readily dissolves oils, many resins, sulphur, phosphorus and other like bodies. Oil of turpentine is used as an internal remedy, and in such cases the redistilled oil should be employed.

Oil of Tobacco.

Synonym.—Oleum tabaci.

This so-called volatile oil is made by the destructive distillation of dry tobacco. It is a dark brown liquid of a strong, peculiar odor. It doubtless contains nicotine and some other peculiar bodies, and the general products of vegetable destructive distillations.

Oil of Valerian.

Synonym.—Oleum valerianæ.

This volatile oil is obtained by aqueous distillation from the root of Valeriana officinalis. (*Lin.*) It is of a yellowish color when recently distilled from fresh valerian, but by age darkens and thickens, acquiring a rank, strong

odor. Oil of valerian contains a hydrocarbon, $C_{10}H_{16}$, valerol, $C_6H_{10}O$, and valeren (borneen), $C_{10}H_{18}O$. By the gradual decomposition of oil of valerian, valerianic acid as one product is produced, and at one time valerianic acid was made from valerian root. At present, however, it is made from fusil oil.

Oil of Wintergreen.

Synonyms.—Oil of gaultheria. Oleum gaultheriæ.

This is made by aqueous distillation of wintergreen herb. It is colorless when fresh and pure, but changes to reddish by age, especially if in tin or iron containers. It is remarkable for its specific gravity, being heavier than water (1.17 to 1.18), therefore, the pure oil will sink in water. Oil of wintergreen has a pleasant aromatic odor, is mainly composed of methylsalicylic acid, $CH_3\,C_7H_5O_3$, and when heated with caustic alkali it will form a salicylate of the alkaline metal, methyl alcohol being liberated. Besides methylsalicylic acid, wintergreen oil contains about 10 per cent. of a hydrocarbon, $C_{10}H_{16}$, named gaultherilene. Wintergreen oil is extensively adulterated, and to suit the price which the purchaser desires to pay for it. Sassafras oil is generally employed, the necessary specific gravity being given by the admixture of the requisite amount of chloroform. In such cases the chloroform can be separated by distillation at less than 200° F., and the odor of sassafras will be at once apparent, upon mixing the residue with a solution of one-half its weight of caustic potash in four parts of hot water. When oil of wintergreen is poured from the hight of ten or twelve inches in a thin stream, upon the surface of water, it should sink and form a layer at the bottom.

EXAMINATION OF URINE.

It has been found that certain abnormal conditions of the system are accompanied by changes in the chemical constituents of the urine. It is not designed that this department of our book shall include microscopic examinations, etc., but that it shall enable the student and the physician to apply reliable tests for the chemical substances found in urine, and which it is often necessary to consider. Few busy practitioners have the time to devote to quantitative examinations, and few can spend time to experiment regarding the constituents of normal urine.

In consequence of these facts, the tests given shall be such as are adapted to the every-day work of the practitioner or druggist. They shall be clearly and tersely stated, and such abnormal substances considered as upon identification may aid physicians in the treatment of disease.

Normal Urine has usually a specific gravity of from 1.012 to 1.018 at the temperature of 60° F. This is increased or diminished by circumstances, such as diet, exercise, temperature, etc. Therefore, increased or diminished specific gravity is not a certain indication of disease. If, however, the amount of urine voided be large, and the specific gravity from 1.030 to 1.040, the presence of sugar may be suspected, or urea in excess. The urinometer is a very convenient little instrument for taking the specific gravity of urine, and sufficiently accurate for most purposes. In using it, bring the urine to the proper

temperature, and fill the cylindrical glass with urine, so that the urinometer floats. Remove the foam if any be present, then place the glass in a perpendicular position, float the urinometer, and read the scale. The point where the surface of the urine strikes the scale is the specific gravity of the urine. Urinometers are graduated with the last two figures of the decimal of the regular hydrometer. Thus 1.000 of the specific gravity hydrometer is represented by 00 upon the urinometer, 1.010 by 10, 1.018 by 18, etc. Instead of normal urine being represented by from 1.012 to 1.018, it will be from 12 to 18 by the urinometer.

Albumen.—Permit the freshly voided urine, if turbid, to stand quietly an hour in a cool location. If it has not then settled so that a clear solution can be decanted from the sediment, filter it. Fill a test-tube about one-fourth or one-third with the clear liquid, and test it with blue litmus paper. If it be not acid, add to it acetic acid until a distinct acid reaction results, care being taken to avoid an excess of acid. If the urine is naturally acid in reaction, the addition of an acid is unnecessary. Bring the mixture to a boil. If it remain transparent, albumen is absent; if it becomes cloudy or filled with precipitate, either albumen or earthy phosphates are present. Add a few drops of nitric acid, and if the mixture remains cloudy, or the precipitate does not dissolve, *albumen* is present; if the liquid becomes transparent, the precipitate or the turbidness resulted from the presence of *earthy phosphates.*

Earthy Phosphates.—The foregoing test is reliable, if they are in large amount. To examine directly for earthy phosphates, add an excess of ammonia water (ammonium hydroxide) to the clear urine. Boil a portion in a test-tube; if a precipitate follow, *earthy phosphates* are present.

Sugar (*Glucose.*)—This is identified as follows: Mix one part by measure of Fehling's soda solution with one part by measure of Fehling's copper solution. Place

this to the depth of half an inch in a test-tube, and mix with it as much clear urine as the reagent measures. Bring the mixture to a boil over an alcohol lamp; if a precipitate results (cuprous oxide, Cu_2O), ranging in color from yellow to brownish-red, *sugar* was present in the urine.

Sometimes the mixture throws down upon boiling, a heavy bluish precipitate, which may tend to cover the color of traces of cuprous oxide. In such cases, mix the Fehling's solution with the urine as before directed, and place the mixture in a cool situation. After twenty-four hours a red precipitate will have fallen, if sugar was present in the urine. Fehling's solutions † are to be made as follows:—

Soda solution.—Dissolve 2681 parts of neutral potassium tartrate in 9300 parts of solution of caustić soda, specific gravity 1.12.

Copper solution.—Dissolve 536 parts of pure crystallized sulphate of copper in sufficient distilled water to bring to the measure of the "soda solution."

Urea.—This principle is always present in urine. It results from the destruction of the nitrogenous tissues of the body, and is the means by which nitrogen is eliminated. In health, the daily excretion amounts to about an ounce during twenty-four hours. An animal diet increases the amount of urea, a low diet diminishes it. Certain diseases are accompanied by increased disintegration of the nitrogenous principles of the body, and in such cases an abnormal amount of urea is excreted.

Excess of Urea.—Mix in a test tube one measure of officinal nitric acid with two measures of the clear urine, which has not been permitted to evaporate by exposure to the atmosphere. Cork the tube, cool, if necessary, until the temperature is reduced to 60° F. (15.5° C.) and permit it to rest quietly half an hour. If the mixture be a mass of crystals, or contain crystalline tufts, urea was present in the urine in abnormal proportion. This test is based

† These solutions are to be kept separate, and mixed as required, and just before each experiment.

upon the fact that nitrate of urea is less soluble than urea. The relative proportion of water and urea of normal urine is such that the nitrate of urea formed by the addition of nitric acid remains in solution. Sometimes there is a deficiency of water, thus making the urea above the proper proportion, while in reality the urea is only excreted from the system in normal amount. To determine this point, measure the urine which passes during twenty-four hours, and bring it to the bulk of 40 fluid ounces by addition of distilled water, then apply the test previously given for excess of urea. A crystalline formation will show excess of urea.

Pus.—Permit the urine to settle. Decant the clear liquid and add to the sediment a solution of potassium hydroxide (caustic potash). If pus be present a gelatinous mass results; if pus is found in the sediment, albumen may be expected in the clear liquid previously decanted.

Bile Coloring Matter—Pour hydrochloric acid into a beaker glass to the depth of half an inch. Stir urine into this until the acid is colored, then through a funnel tube introduce a layer of nitric acid directly upon the bottom of the beaker. If bile coloring matter is present a play of colors is exhibited where the liquids are in contact. Mix the liquids with a glass rod, color-changes are exhibited throughout the entire liquid.

Blood.—Place a specimen of urine, as voided, in a test-tube and add solution of potassium hydroxide (caustic potash), then warm gently. A precipitate will follow (earthy phosphates) which will be blood-red if blood is present. Otherwise the precipitate will be white.

If the urine should be alkaline in reaction, add first a small amount of solution of magnesium sulphate, then the solution of caustic potash. The microscope and spectroscope are valuable in determining the presence of blood.

Uric Acid.—This is present in the urine of carnivora. It occurs also in combination, forming urates. Uric acid is never present to form a sediment in freshly voided normal urine.

Test.—Uric acid in abnormal amount precipitates, generally brick-dust colored. Place a little of the sediment in a porcelain dish with a small portion of water, then add a few drops of nitric acid, warming carefully. When the uric acid has dissolved evaporate the solution almost to dryness on a water bath. Add now a drop of ammonia water, or hold over the surface of the mixture a glass rod moistened with ammonia water. A beautiful purple-red color results (murexide) if uric acid was present in the original precipitate.

TABLE OF WEIGHTS.

480.0	grains Troy	equal	1	ounce Troy.
437.5	"	"	1	" Avoirdupois.
7000.0	"	"	1	pound Avoirdupois.
5760.0	"	"	1	" Troy.

	English Grains.	Troy Ounces.
Milligramme	0.015432	0.000032
Centigramme	0.154323	0.000322
Decigramme	1.543235	0.003215
Gramme	15.432349	0.032151
Decagramme	154.323488	0.321507
Hectogramme	1543.234880	3.215073
Kilogramme	15432.348800	32.150727
Myriogramme	154323.488000	321.507267

1 Grain equals 0.064798950 Gramme.
1 Troy Ounce equals 31.103496 Grammes.
1 Pound Avoirdupois equals 0.45359265 Kilogramme.
1 Cubic Centimetre of water weighs 1 Gramme.

RELATION OF WEIGHT TO MEASURE.

(TEMPERATURE 60° F.)

1 Minim	of distilled water	equals	0.9493	Grains.
1 Fluidrachm	"	"	56.9618	"
1 Fluidounce	"	"	455.6944	"
1 Pint	"	"	7291.1107	"
1 Gallon	"	"	58328.8862	"

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